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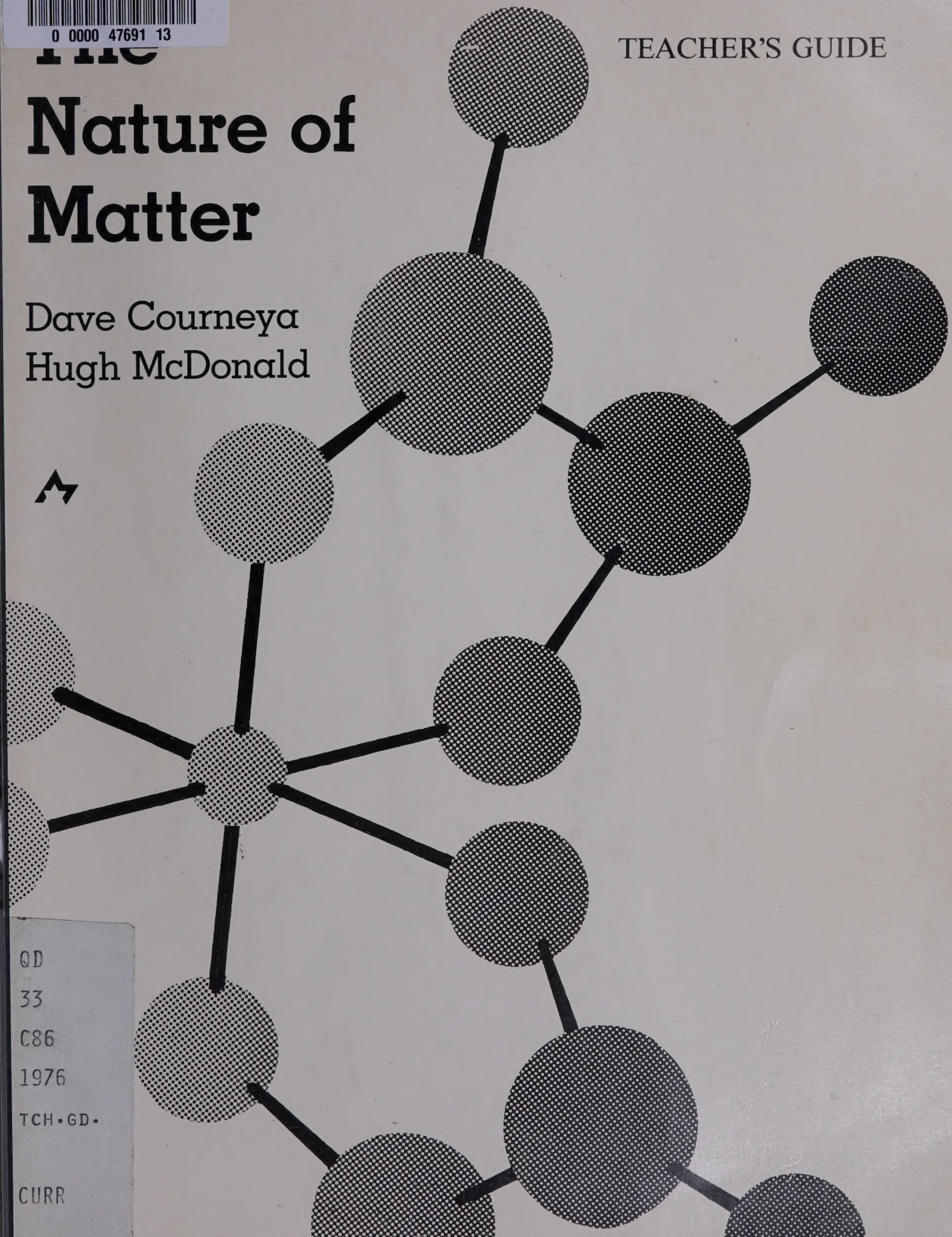
TEACHER'S GUIDE

The Nature of Matter

Dave Courneya
Hugh McDonald



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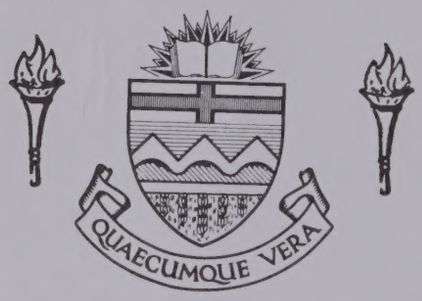

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Dave Courneya

Assistant Head of Science
East Northumberland Secondary School

Hugh McDonald

Head of Science
East Northumberland Secondary School

D.C. Heath Canada Ltd. Toronto

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Foreword

It is not our intention to write a long dissertation on the nature of matter. Our intention in this *Teacher's Guide* is to try and provide some useful suggestions on how we think you can make the best use of the text. Before we go too far, however, we should make our motives clear to you.

Course Objectives

We recognize that aims and objectives can be very pedantic, but we nevertheless feel that the teacher must have some clear idea of what he or she is trying to do in the classroom.

It seems to us that the most important single thing we can do for students is to teach them how to handle information. To be more specific, we must teach them how to:

1. gather
2. organize
3. evaluate
4. interpret

} data or information.

We feel that every single citizen of this country or any other country would benefit from this goal. Each part of this process is valid and necessary and no one part should take precedence over any other. In addition, some of our other aims, goals, aspirations and objectives are as follows.

1. A science student should learn to go from the concrete to the abstract; from observations to conclusions to explanations.
2. A science student should learn that there may be several explanations for a given situation and that he *should* select the best of these explanations on the basis of the evidence available.
3. A science student should gradually change from being qualitative to quantitative in his thinking.
4. A science student should learn to control the direction of his thinking. He should learn to "discover" relationships or ideas from experimental data.
5. A science student should become a person who asks himself "why?".

In trying to reach these goals a student enrolled in a science program should:

1. increase his knowledge of nature, of matter, of energy and of life and the relationships that exist among them;
2. increase his ability to follow instructions and carry out a laboratory-oriented exercise;
3. develop an interest in science and an awareness of the role of science in the modern world;
4. become more organized in his approach to a problem; learn that observation comes before conclusion; that explanations come after conclusions;
5. learn to ask himself questions that relate to a specific problem and to question the answers to preceding questions; learn that questions always precede answers;
6. learn that any question (well, many questions) can be answered in the same way that science questions can be answered and that a rational or logical approach to a problem is frequently better than an emotional approach.

Finally, the student who completes a science program must be able to demonstrate by manual, verbal or written response, significant improvement in his ability to carry out the following (points are *not* listed in order of importance):

1. identify certain items of scientific equipment;
2. read and carry out a set of sequential instructions related to a laboratory exercise;
3. perform certain routine skills without the aid of written instructions, i.e., lighting a burner;
4. extract the observations that are relevant to the problem under investigation;
5. compile a written record of what transpired during the laboratory exercise;
6. separate fact from non-fact;
7. organize the relevant data into a more meaningful form, i.e., graph, table, diagram;
8. interpret the relevant data by:
   (a) drawing a conclusion,
   (b) offering an explanation,
   (c) developing an hypothesis,

(d) suggesting a generalization,
(e) finding a regularity;
9. develop an experiment that will test a hypothesis;
10. state with reasons the limitations of any laboratory exercise;
11. apply knowledge gained in previous investigations to a similar situation;
12. solve problems based on scientific principles or laws;
13. write a concise and clearly worded report based on a laboratory exercise or other research topic;
14. use the vocabulary of science correctly;
15. give examples of the application of scientific principles to the world of today;
16. derive or extract information from related materials of a scientific nature;
17. assimilate and recall as required an appropriate portion of the accumulated body of scientific knowledge;
18. recognize that science is a human activity;
19. spell, write a complete sentence, use proper grammar and use a dictionary;
20. follow laboratory safety procedures;
21. follow good "housekeeping" procedures.

Evaluation

It is our belief that evaluation should vary as much as possible in an attempt to determine if the objectives are being achieved. We attempt to evaluate by including all of the following:

1. examinations
2. class tests
3. laboratory reports
4. laboratory tests
5. research topics
6. homework
7. book reports
8. laboratory skills
9. problem solving
10. essays

We must emphasize how important we think the laboratory-test evaluation can be. As we go through this guide we will try to be specific on evaluation.

Text Contents

Having cleared the air somewhat perhaps we could get to more practical matters. *The Nature of Matter* contains nine units, but we do not see how anyone could cover all of them in a single year. It is our feeling that the first five units should be considered as core material and the rest as optional. In the first five units we have tried to cover the important concepts that are fundamental to a first chemistry course.

Unit 1 provides an introduction to chemistry. It begins with the processes of science and then looks at matter and the changes that matter can undergo. The ideas of physical and chemical change are reviewed.

Unit 2 deals with what we might call the molecular model. We look at the scientific evidence that has led us to believe in molecules. The important ideas established here are the fundamental laws of chemistry, i.e., conservation of mass, definite proportions, Boyle's Law, Charles' Law, Avogadro's Hypothesis and the Ideal Gas Law.

Unit 3 deals with the atomic model. Again we have tried to offer the scientific evidence that has led us to believe in the idea of sub-atomic particles and their arrangement inside the atom. We take the student up to the Bohr atom and the idea of the stable octet proposed by Lewis. This enables us to establish valence, elementary bonding, formula, equations and conductivity.

Unit 4 uses descriptive chemistry to convey the basic ideas of stoiciometry. Included here is the work on acids, bases and neutralization, molarity and molar solutions, titrations, oxidation and reduction and equations.

Unit 5 brings much of the first four units together by focusing on the experimental evidence that led to the organization of the periodic table.

Unit 6 focuses on the unique properties of carbon and its compounds.

Unit 7 deals with some substances that students should be familiar with.

Unit 8 contains laboratory skills that a student should be familiar with by the time he has finished the course.

Unit 9 contains some ideas for interesting independent study projects.

Laboratory Safety

We could say a great deal about lab safety, but most of it is common sense. What we will do is list the

basic pieces of equipment that a lab should have if it is to be relatively safe. Most school labs lack at least one of these items; many schools lack several of them. If your school falls into this category, we suggest you make a written request to your principal to have your lab brought up to these very minimal standards. We have a double-barrelled motive for saying this.

1. You may get the items you need, thereby making the lab safer for your students.
2. If anything does happen to a student, it can at least be shown that you tried to make your lab safe.

Safety Equipment

You must have the following:
safety glasses for every student
rubber bands for tying hair back
fire extinguishers: CO_2, dry chemical and sand pail
fire blanket
eye wash station
first aid kit
fume hood
vented storage area
lockable storage area
broom and dustpan

After you have discussed lab safety, have your students look at the photo on page 27 of the text. Which safety procedures are being obeyed? Which are being violated?

Equipment and Supplies

We recommend that you set your lab up so there are permanent stations spread around the room. The following equipment should be kept at each of these stations:

Item	*Quantity*
test tubes with a rack	assortment
beakers: 1000 ml	1
600 ml	1
400 ml	1
250 ml	2
150 ml	2
beaker tongs	1
bunsen burner	1
crucible tongs	1
dropper and bulb	1
Erlenmyers (250 ml)	2
gas collection: bottles or jars	4
gas igniter	1
graduates 100 ml	1
10 ml	1
pneumatic trough	1
retort stand and ring	1
spatula or scoopula	1
test tube clamps	2
test tube holder	1
wire gauze with asbestos centre	1

The following equipment should be available but not kept at the student stations except when it is being used. Amounts suggested can accomodate sixteen groups, with a small allowance for loss or breakage.

Item	*Quantity*
asbestos squares (25 cm × 25 cm)	20
balances (centigram)	2 automatics or 16 manual
balloons	many
"black boxes"	20
burettes (50 ml)	20
carbon electrodes	20
cathode ray tubes (borrow from physics)	
clay triangles	20
conductivity testers	20
crucibles and lids	40
cylinders: 1000 ml	6
250 ml plastic	20
25 ml	20
deflagrating spoons	20
density vials	20
drying tubes	20
distilled water	lots
ebonite rods (physics)	
electrical leads (physics)	
eudiometer tubes	2
evaporating dishes	20
flame spreaders	20
flasks (side-arm)	20
funnels	20
gas discharge tubes (physics)	
gas measuring tube (50 ml)	2
gas law tubes (see pages 23, 24)	20
Geiger counter (physics)	
glass beads	5 lbs.
glass plates (10 cm × 10 cm)	80
glass rods	20
glass tubing	assortment
glass tubing (11 mm × 120 cm)	20
Hoffman apparatus	2
hot plate	2
induction coil or Tesla coil	1

knives	2
light sources	2
manometers	20
mortar and pestle	20
pipettes and bulbs	20
platinum wires in handles	20
power supplies (6v) (physics)	
rubber and cork stoppers	assortment
rubber tubing	assortment
serum caps	50
spectrometer (hand)	20
syringes (10 ml disposable)	20
thermometers (–10 + 120°C)	20
thistle tubes	20
timers (physics)	
tin cans: 48 oz.	20
10 oz.	20
tweezers	20
U-tubes	20
volumetric flasks: 1000 ml	6
100 ml	20
wash bottles	20
watch glasses	40

Chemicals and Software

The quantities specified on these pages are minimum quantities required for a single class. In some cases you will have lots left over the next year.

Item	*Quantity*
Stock Acids and Bases	
acetic, glacial	1 × 5 lbs.
ammonia, aqueous (conc.)	1 × 5 lbs.
barium hydroxide	1 kg
benzoic acid	500 g
calcium hydroxide	1 kg
hydrochloric acid (conc.)	2 × 6 lbs.
nitric acid (conc.)	1 × 6 lbs.
potassium hydroxide	1 kg
sodium hydroxide	2 kg
sulphuric acid (conc.)	1 × 10 lbs.

Elements (where amounts not specified, small quantity only)	
aluminum	large pieces
antimony powder	
argon (in a discharge tube)	
boron	
bromine	
calcium (metal turnings)	500 g
carbon: charcoal	
coke	
graphite	
lampblack	
large lumps	
soft coal	
chlorine gas	lecture cylinder
copper: foil	100 g
large pieces	
shot	500 g
hydrogen gas	lecture cylinder
iodine	500 g
iron: granular	500 g
large pieces	
lead metal: foil	500 g
large pieces	
shot	500 g
lithium metal	100 g
magnesium ribbon	100 g
neon gas (inside a discharge tube)	
oxygen gas	lecture cylinder
potassium metal	100 g
silicon	100 g
silver	
sodium metal	100 g
sulphur: lump	500 g
sublimed	500 g
zinc: granular	500 g
mossy	1 kg
technical	1 kg

Chemicals and Other Supplies	*Quantity*
acetone	2 pts.
aluminum foil	1 pkg.
aluminum oxide	100 g
ammonium nitrate	500 g
antacid tablets	assortment
asbestos wool	100 g
benzene	1 pt.
brom thymol blue	5 g
calcium carbide	100 g
calcium carbonate: chips	500 g
powder	500 g
calcium chloride anhydrous	500 g
calcium oxide	500 g
carbon tetrachloride	2 pts.
cloth (cotton)	
cobalt chloride	500 g
cobalt chloride paper	12 vials
coconut oil	500 g
copper carbonate	500 g

copper sulphate	500 g	pins (T-large)	4 doz.
copper wire: 18 gauge	1 roll	Plasticine	3 lbs.
0.05 cm	1 roll	potassium carbonate	500 g
cotton swabs	1 pkg.	potassium chlorate	1 kg
crude oil	2 pts.	potassium chloride	500 g
cyclo hexane	1 pt.	potassium permanganate	100 g
cyclo hexene	1 pt.	sebacoyl chloride	100 g
detergent	500 g	silicon oxide	100 g
dialysis tubing	10 m	silver nitrate	100 g
dry ice	1 kg	soap	500 g
ethanol	2 pts.	sodium acetate	100 g
filter paper	12 pkg.	sodium bicarbonate	500 g
food colouring		sodium bromide	100 g
formaldehyde (40%)	1 pt.	sodium carbonate	500 g
glucose	100 g	sodium chloride	500 g
glycerol	1 pt.	sodium iodide	100 g
hexane	1 pt.	sodium oxide	100 g
1-6-hexane diamine	100 g	solder	roll
hydrogen peroxide (6%)	2 pts.	splints (wood)	gross
magnesium oxide	100 g	steel wool	pkg.
manganese dioxide	500 g	straws (drinking)	gross
methanol	2 pts.	sugar	500 g
mineral oil	2 pts.	thread	spool
paper (litmus pink, blue)	2 doz. ea.	toluene	1 pt.
paper matches	gross	urea	100 g
paraffin candles	2 doz.	wood shavings	
phenol	1 pt.	yeast	12 pkg.
o-phthalic acid	100 g		

Solutions

The following solutions are required in the concentrations indicated. We have included the amount of solute required to prepare 1 ℓ of 1M solution. It should be scaled up or down an appropriate amount for other concentrations. In most cases the quantities for solids are given in the anhydrous form. For acids we have started from the commercial stock available.

Solution	*Formula*	*Concentrations M*	*Solute per 1 ℓ of 1 M Solution*
acetic acid	CH_3COOH	0.1, 0.5, 1, 5	59 ml of 17 M/ℓ
ammonium chloride	NH_4Cl	6	53.5 g
ammonium hydroxide	NH_4OH	0.1, 1, 6	65.5 ml 15 M/ℓ
barium chloride	$BaCl_2 \cdot 2H_2O$	0.1	244 g
barium hydroxide	$Ba(OH)_2$	sat.	About 100 g/ℓ will saturate.
brom thymol blue	—	—	1 g dye dissolved in 16 ml 0.1 M NaOH dilute to 1 ℓ
calcium chloride	$CaCl_2$	0.1	111 g
calcium hydroxide	$Ca(OH)_2$	sat.	About 2 g/ℓ will saturate.
copper II bromide	$CuBr_2$	0.1, 0.5	223 g
copper II chloride	$CuCl_2$	0.1	134 g

Solution	*Formula*	*Concentrations M*	*Solute per 1 ℓ of 1 M Solution*
formic acid	HCOOH	0.1	46 g
hydrochloric acid	HCl	0.1, 0.5, 1, 2, 4, 6, 8	83 ml 12 M/ℓ
hydrogen peroxide	H_2O_2	3%	stock 6%
hypochlorous acid	HClO	1 M	52.5 g
iron III chloride	$FeCl_3 \cdot 2H_2O$	0.2	170.5 g
iron III nitrate	$Fe(NO_3)_3 \cdot 9H_2O$	0.1	404 g
lead nitrate	$Pb(NO_3)_2$	0.1	331 g
mercury I nitrate	$HgNO_3$	0.1	263 g
nitric acid	HNO_3	0.1	62 ml 16 M/ℓ
phenolphthalein	—	—	1 g of dye dissolved in 600 ml ethanol diluted with water to 1 ℓ
phosphoric acid	H_3PO_4	1 M	98 g
potassium bromide	KBr	0.1	119 g
potassium chlorate	$KClO_3$	0.1	122 g
potassium chloride	KCl	0.1	74.5 g
potassium chromate	K_2CrO_4	0.1	194 g
potassium hydroxide	KOH	0.1	56 g
potassium nitrate	KNO_3	0.1	101 g
potassium permanganate	$KMnO_4$	0.01	158 g
potassium thiocyanate	KSCN	0.002	97 g
silver nitrate	$AgNO_3$	0.1	170 g
sodium bromide	NaBr	0.1	103 g
sodium chloride	NaCl	0.1	58 g
sodium hydrogen phosphate	Na_2HPO_4	6 M	142 g
sodium hydroxide	NaOH	0.1, 1, 0.5, 5	40 g
sodium hypochlorite	NaOCl	5.75% (bleach)	Javex bleach
sodium iodide	NaI	0.1	150 g
sodium phosphate	Na_3PO_4	0.1	164 g
sodium sulphide	Na_2S	0.1	78 g
sulphuric acid	H_2SO_4	0.1, 0.5, 1, 4, 6	55.5 ml of 18 M/ℓ

Unit 1
An Introduction to Chemistry

If you have not yet done so, have the students read the foreword to the student in the text. Spend a few moments the first day showing them where to find the various pieces of safety equipment and explaining some of the safety rules you want them to follow. Spend a few moments going over the equipment illustrated on pages 12 and 13. Point out to them that they are responsible for knowing the names of all the pieces of equipment they will use over the year.

Discuss with them what science is and how scientists operate. Establish the idea that we do labs to gather information; that the analysis of the information sometimes helps us to understand things; that frequently we can suggest alternate explanations for the same data; and that we often come up with as many questions as we do answers.

For homework ask them to read over the first experiment and prepare a data sheet. Don't allow anyone to go to the lab without a data sheet. This is a good practice to follow on every lab exercise.

1.1 Candles

Most students will have no difficulty doing this lab. In the discussion that follows the lab you should be able to review the difference between qualitative and quantitative observations. Obtain from the students a list of the changes that took place in this experiment. Ask for explanations of why these changes took place, and concentrate on the loss of mass and the decrease in height. Ask them to propose a theory that will explain these two changes and still be consistent with the other known facts that were established.

Take a moment to show them how they have been handling the information they gathered in the first part of the lab. Emphasize that "the lab" is more than the gathering of the data – it includes the analysis as well.

You should be able to get some students to propose the phlogiston theory. Have the students read about the theory on page 21 and then discuss it as a possibility that explains what they observed. Prepare them for an experiment in which they will try to catch some of this phlogiston so they may study it. Have them prelab 1.2 for the next day.

1.2 More Candles

Provide as many different sizes of beakers as you can. It can be interesting to use some tall-form beakers and some short-form beakers. Once the student has gathered the data, two things should be apparent:

1. Different beakers give different "going-out" times.
2. Different candles in the same beaker give different going-out times.

Ask for someone to relate these new facts to the ideas established in the first experiment. Does the phlogiston theory still apply here? Ask for other explanations. Ask for explanations as to why the candle goes out when it is covered. Point out that since most beakers have a lip on them the beaker does not actually provide a perfect seal on the supporting surface.

Look at the "Conclusions" section on page 23. Discuss what a conclusion is and how we arrive at one. Be sure that you emphasize that conclusions may not always be "right."

Prepare your students for the next lab by saying that we need more information to help us decide which of the proposed explanations is the best one. If a student suggests an experiment not listed in 1.3, try to make every effort to let him do it.

1.3 Why Does a Candle. . .?

Give the class two consecutive periods to do these mini-labs. Spend your time going from group to group asking what they have learned so far. At the end of the lab ask for some sort of report. We suggest that it not be a formal lab report. (We ask students to hand in the answer to question 6 on page 27 and take up the question the next day.) Be sure to come back to the original question: why does the

candle go out? Consider the explanations in light of the new information.

If possible, show the film loop "The Mouse and the Candle," by Popular Science Publishing (ESA-4). It fits well here and students find it interesting.

Look at the problems on page 27.

Summarize the concepts of observations, lab safety and conclusions or interpretations.

Prepare your students for the next lab by saying that you want to see how heat affects other substances.

In order to explain the burning of the candle one usually assumes that the solid wax melts and collects in the bowl of the candle. In point of fact this is probably an assumption on the part of most students looking at this for the first time. Before you begin experiment 1.4 you should try to make your students question what the substance in the bowl of a burning candle actually is.

1.4a/1.4b Heating Solids

Have your students supply the juice-can lids for this experiment.

Two ideas can be established from this lab:

1. Some substances melt when heated.
2. Substances which do melt when heated do so at a unique temperature.

From this we might be tempted to conclude that substances which melt at different temperatures are different and substances which melt at the same temperature are the same – providing all other properties are the same as well.

In 1.4b students compare the properties of candle material and the liquid from the bowl of the candle. If the properties are the same they will probably conclude that these are in fact the same substance. Don't be alarmed if they do not come to this conclusion. The real point of the exercise is to get students to realize that we identify substances by matching properties.

This is probably a good place to give a test. What follows are a few questions that you might find useful in preparing the test. This is a multiple-choice test but not all our tests are of this type. We would like to express our thanks to many people for suggesting these questions.

Note on Unit Tests

All unit tests appear in a larger type size so they may be reproduced for classroom use.

Unit 1: Sample Test 1

1. To make a candle that will have a larger flame than those used in class you should
   (a) use a thicker wick
   (b) use a longer wick
   (c) increase the diameter of the candle
   (d) increase the length of the candle

2. Which of the following pieces of information about a burning candle were not clearly demonstrable in the experiments you performed?
   (a) It gives off water.
   (b) It uses up oxygen.
   (c) It gives off carbon dioxide.
   (d) It gives off heat and light.

3. One of the primary requirements for a candle is that it
   (a) be made of candle wax
   (b) have a string wick
   (c) burn
   (d) be white

4. A copper coil placed over a candle flame puts the flame out because
   (a) air can no longer get to the wick
   (b) heat needed to vaporize the wax is conducted away
   (c) gases produced can no longer escape
   (d) moisture condenses on the cold copper and puts out the flame

5. When a candle is blown out, a thin line of "smoke" can be seen coming from the wick. This smoke is most likely
   (a) oxygen
   (b) carbon dioxide
   (c) paraffin vapour
   (d) none of these
   (e) all of these

6. From the melting of wax in warm water we learned that
   (a) different forms of the same substance act the same way
   (b) there is probably no change in the wax as it melts and collects in the candle depression
   (c) a lot of hot water is required to melt a little wax
   (d) wax will always melt under high heat

7. One of the following is *not* a qualitative observation.
   (a) The candle is malleable.
   (b) The flame is blue at the base.
   (c) The wick goes completely through the candle.
   (d) The candle grows shorter as it burns.

8. While doing the experiment with different materials on the tin lid, student A found the melting order to be wax, sulphur, tin and lead. Student B, on the other hand, found the order to be wax, sulphur, lead, then tin. Which observations are correct?
   (a) A
   (b) B
   (c) both A & B
   (d) neither A nor B

9. Which of the following is a quantitative property?
   (a) A candle is cylindrical.
   (b) A candle is made of a white greasy material.
   (c) The length of a candle decreases as it burns.
   (d) The candle was 15 cm long.

10. An eccentric science teacher grows tropical plants in his living room and finds that they thrive best in a humid atmosphere. Which of the following sources of heat would suit his needs best? (Assume that all four can supply heat at the required rate.)
   (a) 30 electric toasters
   (b) a car radiator with hot water circulated through it
   (c) 10 electric hairdriers
   (d) 6 bunsen burners

11. The statement, "The candle gives off light and heat as it burns." is
   (a) an interpretation
   (b) an observation
   (c) a qualitative observation
   (d) a quantitative observation

12. A good experiment is one that
   (a) answers all the questions you originally had to your complete satisfaction
   (b) proves one or the other of the postulated hypotheses
   (c) does none of the above
   (d) answers some of your questions, but raises new additional questions

13. When a piece of aluminum foil was inserted around the wick below the flame, the flame went out because
   (a) the aluminum foil conducted sufficient heat away from the flame to lower the temperature below the melting point of the wax
   (b) the aluminum foil prevented the melted wax from flowing up the wick by capillary action
   (c) the aluminum foil prevented oxygen from getting to the flame
   (d) the aluminum foil absorbed the melted wax like a piece of blotting paper

14. Which of the following statements is correct? "The dark zone of a candle flame is...
    (a) composed of non-combustible vapours
    (b) an invisible flame
    (c) composed (in part) of vaporized wax
    (d) no different than the rest of the flame
    (e) none of the above

15. Which of the following statements is not an observation?
    (a) A colourless liquid forms in the bowl near the base of the wick.
    (b) Sulphur is a yellow solid.
    (c) Carbon is formed on the wick of the burning candle.
    (d) A fog or mist appears on the inner surface of the beaker.
    (e) Lead changes into a liquid when heated strongly.

The next three sections show students some of the important considerations one must make when looking at data.

Conditions that Matter

Students must be aware that the conditions under which the data is gathered are important. Changing the conditions may change the outcome of an experiment. Students should learn to consider these conditions and to control them where possible. Point out to them that when one designs an experiment, one must try to specify and control all the conditions that could affect the outcome of the experiment. Point out as well that when doing a lab one should follow the procedure outlined as closely as possible in order to duplicate the conditions outlined in the procedure.

Interpretations

The purpose of this exercise is to show that one must gather all the data possible before drawing any conclusions: i.e., keep an open mind until all the evidence is in.

Hypothesizing

The main idea to be established by this exercise is: a conclusion is only as good as the data that produced it. Conclusions must always be modified to suit new information. Conclusions are not "true," they are just the best we can do with the information we have. Spend some time pointing out that this is what science is all about – the "growth" of the model to include new information.

Physical Change/States of Matter

These two sections are a review of previous work. Two experiments on sublimation are included, but we suggest that you do only one of them. It may even be possible to cover these sections by the Socratic method.

1.6 Other Changes

Encourage your students to answer the questions for each part before going on to the next. The object is to see that all changes are not physical and to call the new change a chemical change. We suggest that you define physical and chemical change in terms of the reactants and the products.

In a physical change, the product of the change is the same substance as the material we started with. (Refer back to lab 1.4a, where we decided that we could identify a substance by its set of properties.)

In a chemical change, the product of the change is not the same substance as the material we started with.

Having seen something of what energy does to a substance, you can set the stage for determining the "how much" for physical and chemical change.

Prepare the student for the calculations by going over the "Calorimetry" section, which begins on page 42. (It has been pointed out to us that while we have used the calorie for many years, it is not a unit of the English system, as we stated in the text. Our apologies!)

1.7a Heat Released by Burning Candle Material/1.7b Solidifying Candle Material

After trying several different approaches to this two-part lab we have finally settled on a four-day time period.

Students do the lab and try the calculations on their own the first day.

We go through the calculations with them and get the results of each group's calculations on the board the second day. On the same day, discuss those results, determine the class average and estimate the amount of error.

On the third day, students do the second part and try the calculations on their own.

On the fourth day, we discuss the results, determine the average values for the whole class, estimate the errors and compare with the previous lab results.

For evaluation purposes we normally have groups of two, so one partner hands in his attempt at the calculation for the first part and the other partner hands in his attempt at the second part.

Overall our objective is for the student to see that the amount of energy associated with a physical change is much smaller than the energy associated with a chemical change. The experiment shows this very well in spite of the rather large errors involved.

The student should now be ready to tackle the problems on page 48.

Once you have finished the problems based on calorimetry, you will probably want to give a single-problem test. We would allow 15 min to do the following problem (slide rules and calculators allowed).

Problem

Two chemistry teachers wishing to make a cup of tea in the laboratory (Never lick the spoon!) find that the gas supply has been turned off and they cannot

use the bunsen burner to heat the water. They decide to use a pocket butane lighter instead. Suddenly they realize that while they are making tea they can also determine the heat of combustion of butane while they are waiting for the water to heat. They proceed in a manner similar to experiment 1.7a and collect the following data:
mass of butane lighter = 122.00 g
volume of water heated = 206.3 ml
mass of water heated = 200.00 g
initial temp. of water = 20°C
temp. of water after 3 min of heating = 31°C
temp. of water after 7 min of heating = 52°C
temp. of water after 15 min of heating = 100°C
mass of lighter after 15 min of heating = 120.40 g

Use this data to estimate the heat of combustion of butane. Be sure to show all your calculations and any assumptions that you make.

The remainder of Unit 1 concerns itself with mixtures, colloids and solutions. For some students this may be mostly review material and you will want to pass through it quickly. If this is the case, we suggest that you do as we do and allow each student to progress at his own pace, doing only those labs he feels will aid his understanding. We allow five periods of class time for this.

For those who wish to follow a more conventional path, three important concepts must emerge.

1. Pure substances are almost never found in the pure form in nature.
2. One substance can be separated from another by looking for a difference in properties.
3. When two (or more) pure substances are mixed together one of several things may happen:
   (a) a mixture may form;
   (b) a colloid may form;
   (c) a solution may form;
   (d) something else may happen, as we saw in experiment 1.6, Part C – a chemical change may occur.

1.8 Mixing Liquids and Solids

This is more an exercise in observation than anything else. We ask our students to hand in the completed table at the end of the lab period. In the final column of the table the student should state his opinion as to whether the resulting combination is a mixture or a solution.

You should spend some time here discussing mixtures and solutions. This is a good place to introduce the words *homogeneous, heterogeneous, solute, solvent, dissolved* and *phase.* It is also a good time to point out that when a chemist says *substance* he really means *pure substance* and when he says *mixture* he usually means *mechanical mixture.*

1.9 Separating the Constituents of Mechanical Mixtures and Solutions

This experiment requires little comment. Make sure that your students realize in the discussion that in order to separate two (pure) substances we must find a difference in properties on which we can base the separation.

1.10a One and One Makes Two

This experiment looks at an aspect of solutions that has been neglected up to this point in the text. Two important ideas emerge from this.

1. There is no law of conservation of volume. (You can ask students about conservation of mass if you like – that lab comes up a little later.)
2. In order to explain what is happening, some students may wish to invoke the idea of molecules and the "spaces" between them. You should then go back and talk about dissolving in terms of these little particles that are too small to see. This prepares students for the next lab, on colloids.

1.11 Colloids

You may wish to skip this section on colloids. They are not important to a beginning chemistry course. What can come out of this experiment and the ones dealing with mixtures and solutions is the concept that we tend to classify matter according to the size of the particles. Mixtures are composed of particles that are big enough to see. Solutions are composed of particles too small to see. Colloids are somewhere in-between.

You will most likely have to help some of your students to see the cone produced – the so-called Tyndall effect.

1.12 Solubility of Potassium Chlorate and Copper Sulphate

It is quite likely that your students will have covered some of this material in previous courses. You may wish to move through this section quite quickly.

This experiment should give students the idea that the amount of a substance that dissolves depends upon temperature. They should also realize when they compare data that the solubilities of different substances are different.

1.13 Preparing Solubility Curves

This experiment will work out fairly well if you have the students use the average figures for the whole class. Students must work very hard to maintain the temperature that you have given them. We suggest that you use only the following temperatures: 10, 20, 30, 40, 50°C. Assign the same temperature to several groups. Make sure that students have a good look at the data before they start to average the figures. Spend some time talking about conditions that matter and the effect that some of these conditions will have on the results. Be sure to point out the difference between the curve that we get from our data and the solubility curve that we usually see.

In the summary to this unit it may be useful to take this approach: ask the question, "What will happen when two different pure substances are placed together in the same container?" In the discussion that follows point out that there are several possible answers:

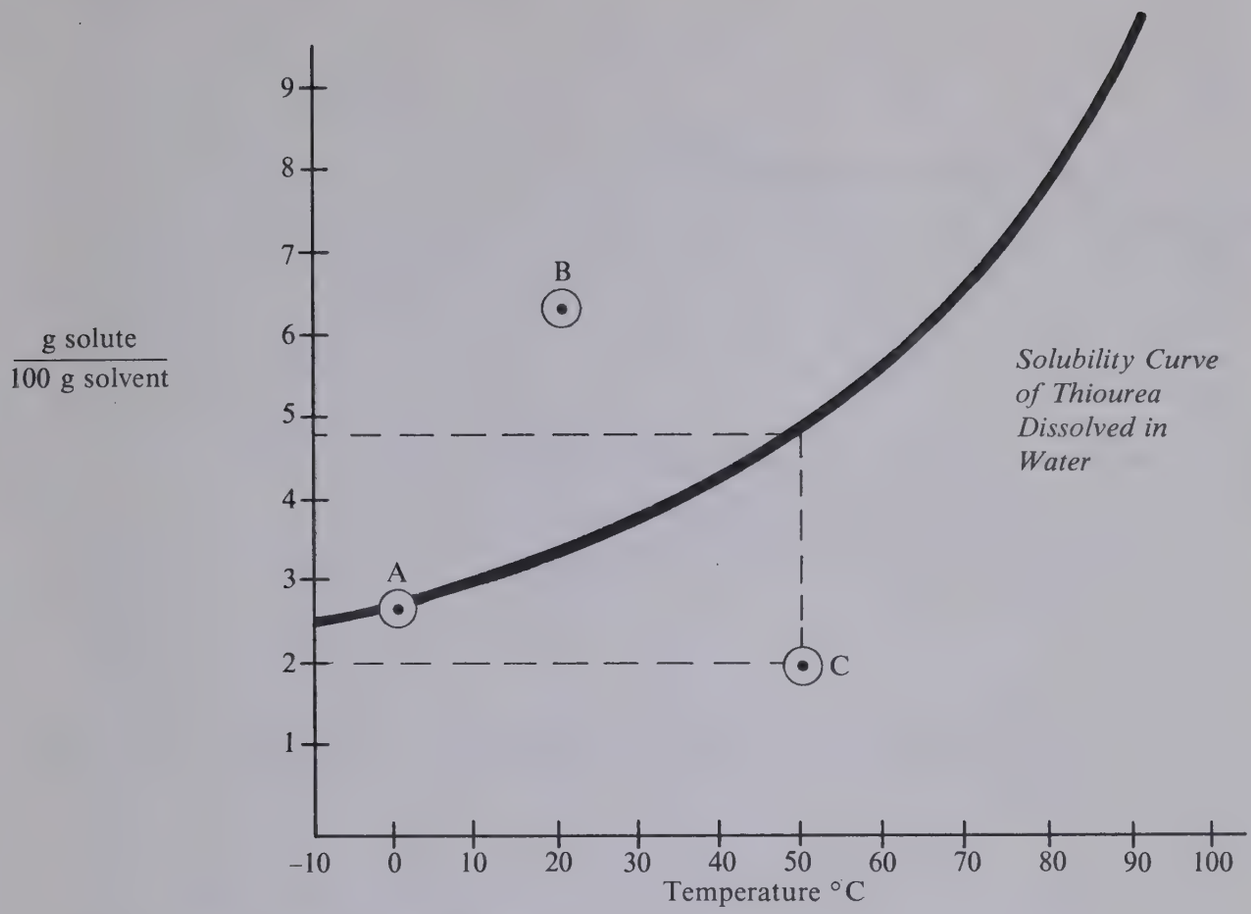
1. a mixture may form;
2. a solution may form;
3. a colloid may form;
4. a chemical reaction may occur.

We would look at the properties of the "product" in order to decide which had happened. The overall theme of this first unit could well have been "properties."

When doing the end-of-chapter questions, point out that we often make the assumption that the density of a solution is 1 g/ml. Sometimes we also assume that the solute occupies no volume. Both are bad assumptions, but they allow us to do some simple calculations from this section.

You might want to use the following to test the section on solutions and mixtures.

Unit 1: Sample Test 2



Solubility Curve of Thiourea Dissolved in Water

1. What type of solution is represented by point A?
   (a) unsaturated
   (b) saturated
   (c) supersaturated
   (d) mechanical mixture
   (e) colloid
2. What would happen if more thiourea crystals were added to solution B?
   (a) It would dissolve.
   (b) It would remain undissolved.
   (c) It would grow larger.
   (d) There would be a chemical reaction.
   (e) It would be exothermic.
3. How much more thiourea could be dissolved in a sample of solution C that contained 50 g of water?
   (a) none
   (b) 4.75 g
   (c) 2.0 g
   (d) 2.75 g
   (e) 1.37 g

4. How many grams of thiourea would dissolve in 100 g of water at 80°C?
   (a) 77 g
   (b) 770 g
   (c) 7700 g
   (d) 7.7 g
   (e) 0.077 g
5. What is the lowest temperature at which 2.5 g of thiourea will completely dissolve in 100 g of water?
   (a) −10°C
   (b) 50°C
   (c) 100°C
   (d) 25°C
   (e) 75°C

The following questions refer to the graph on page 66 of your text.

6. Which of the substances listed is *least* soluble in water at 60°C?
   (a) potassium chlorate
   (b) sodium chloride
   (c) calcium acetate
   (d) ammonium chloride
   (e) barium hydroxide
7. Which of the substances listed shows a decrease in solubility as temperature increases?
   (a) potassium nitrate
   (b) calcium acetate
   (c) barium hydroxide
   (d) sodium chloride
   (e) potassium chlorate
8. The solubility of ammonium chloride at 60°C is
   (a) 4.8 g
   (b) 48 g
   (c) 48 g/100 g water
   (d) 30 g
   (e) 57 g/100 g
9. The substance that is most soluble in hot water is
   (a) barium hydroxide
   (b) potassium nitrate
   (c) ammonium chloride
   (d) sodium chloride
   (e) calcium acetate
10. Solubility is generally expressed as
    (a) g/100 g of water
    (b) kg/ℓ of water
    (c) kg/100 g of water
    (d) g/100 g of solution
    (e) kg/100 ml of water

11. Which of the following is not a property of mechanical mixtures?
    (a) may be separated by filtering
    (b) has two phases
    (c) is homogeneous
    (d) is heterogeneous
    (e) constituents may be mixed in any proportion

12. To separate sugar from a sugar and calcium carbonate mixture, you may take advantage of which of the following?
    (a) differences in particle shape
    (b) differences in colour
    (c) differences in solubility
    (d) effect when mixed with alcohol
    (e) differences in density

13. If 50 ml of oil and 50 ml of water are mixed together, the volume will be
    (a) 100 ml
    (b) more than 100 ml
    (c) less than 100 ml
    (d) about 50 ml
    (e) none of these

14. The most distinctive property of colloids is the fact that
    (a) they are transparent
    (b) they have one phase
    (c) they are homogeneous
    (d) they have a large surface to mass ratio
    (e) light passes through them easily

15. Which of the following is not a solution?
    (a) ginger ale
    (b) granite
    (c) steel
    (d) gasoline
    (e) hard water

16. Which of the following is *not* a mechanical means of separating the constituents of a mixture?
    (a) stirring with a magnet
    (b) panning with water
    (c) filtering
    (d) stirring with water
    (e) straining with a sieve

17. One difference between mixtures and solutions is that mixtures are always
    (a) composed of two different substances
    (b) formed with substances in the same state
    (c) homogeneous
    (d) heterogeneous
    (e) easy to separate

18. In order to separate the components of a solution one could use the properties of
    (a) colour
    (b) density
    (c) differences in boiling points
    (d) chemical composition
    (e) particle size.

19. When dialysis tubing is used to successfully separate two substances, the separation is based on
    (a) colour
    (b) density
    (c) differences in boiling points
    (d) chemical composition
    (e) particle size

20. An amalgam could be described as
    (a) a solid-solid solution
    (b) a pure substance
    (c) a heterogeneous solution
    (d) iron or steel
    (e) a 50-50 mixture of two different metals

21. Air can be classed as
    (a) solution
    (b) mixture of gases
    (c) element
    (d) electrolyte
    (e) compound

22. All of the following concerning a solution are true except
    (a) The homogeneous condition is permanent.
    (b) A solid solution may exist.
    (c) A liquid solution is clear but may have colour.
    (d) Heavy solute particles may settle out.
    (e) Air may be classed as a solution.

23. A chalk and salt mixture could be separated into its components by
    (a) adding water and shaking
    (b) adding water, boiling and filtering
    (c) adding water, filtering and boiling
    (d) adding water, boiling and cooling
    (e) subliming the salt out of the mixture

24. It is difficult to cook potatoes in water at the top of a high mountain because water boils there at a relatively low temperature. This difficulty might be overcome by
    (a) using a larger flame
    (b) using distilled water
    (c) using an aluminum kettle
    (d) using a larger kettle
    (e) putting salt in the water

25. A substance may be considered to be homogeneous when
    (a) every particle is like every other particle
    (b) it is composed of particles whose properties may be quite different
    (c) its density is 1 g/cc
    (d) it is composed of two different elements
    (e) its density approaches that of water

Unit 2
Matter and Gases

Unit 2 has two functions, both vital to the development of this program and to the development of a first-year chemistry course. The primary function of the unit is to provide the students with evidence for the particle nature of matter. In every exercise throughout this unit, this idea should be repeatedly presented to the students. We should continually ask how we can account for the results of each experiment, and, most importantly, if we can account for their outcome in terms of particles.

Laws govern matter, which seems to imply that there is some order in matter. The secondary function of this unit is to expose students to some of the laws that govern matter, and to the way matter behaves under various conditions. By the end of this unit the students should have a reasonable understanding of the following:

1. Law of Conservation of Mass
2. Laws of Definite Proportions and Constant Composition
3. Boyle's Law
4. Charles' Law
5. Avogadro's Hypothesis
6. Law of Partial Pressures

This unit is an exciting exercise in model building, with each experiment either supporting or adding to the particle theory of matter. The climax of the unit is the indisputable fact that equal volumes of gas measured at the same temperature and pressure contain equal numbers of molecules. Students don't want to accept this idea, but all the experimental evidence they have gathered leads to it. The unit is not only exciting, but also requires hard work on your part to constantly keep the students thinking in terms of the particle theory and how each experiment supports or adds to the particle nature of matter.

2.1 Mixing Substances

The intent of this experiment is for the student to see various changes in matter and, more importantly, to ask himself questions about what he has observed but does not understand. We ask the students at the completion of the experiment to hand in not a formal lab report but a list of questions that have arisen from the experiment. Questions that students have asked are:

1. How does ammonium chloride decrease the temperature?
2. What happens to substances like ammonium chloride or sodium hydroxide when they dissolve? They seem to disappear.
3. How can colourless solutions produce a coloured solid?
4. How can a gray substance produce a blue solution?
5. What are the bubbles formed when the acid and the carbonate react?

These are only a few of the questions asked by our students. The teacher can make excellent use of such questions to lead into Unit 2. For example, question 2 discussed with a class will always lead to the idea that such substances must break up into very small particles. And what could be a better answer, since the entire unit is centred around the particle nature of matter? Already the students may be thinking in terms of particles. In each of the experiments that follow in Unit 2, students should be reminded of particles and their explanations directed toward a particle model.

Materials

With careful use, all the materials are relatively safe and readily available in most chemistry lab prep rooms. Note the warning on the use of sodium hydroxide. Working with 1 M HCl is not terribly dangerous, but students should be directed to handle it with special care, and to wash up all spills immediately.

2.2 Mixing Solutions

This experiment, which leads to the Law of Conservation of Mass, produces excellent results. The students should be directed to take accurate measurements and to record them as they find them.

We find that students expect a change in mass when a precipitate forms, and they tend to make a special effort to get some small change in mass. As the students do the experiment we have them record the change in mass for each reaction on the board in a table as illustrated below. The numbers given are actual class data.

Group No.	*Reaction No. 1*	*Reaction No. 2*	*Reaction No. 3*	*Reaction No. 4*	*Reaction No. 5*
1	0.01	0.00	0.00	0.00	0.00
2	0.00	0.00	0.01	0.00	0.01
3	0.00	0.01	0.02	0.01	0.01
4	0.02	0.00	0.02	0.00	0.00
5	0.01	0.02	0.02	0.00	0.00
6	0.01	0.00	0.01	0.01	0.01

Post Lab

In the post-lab discussion, students should be directed to consider how matter can behave in this way. Can the results be explained in terms of particles?

2.3 Heating Substances: Part A

This experiment produces very good results as long as certain precautions are taken. The upper test tube must be tilted mouth downward, but only very slightly, about 1 degree off horizontal. The illustration in the text seems to imply a much greater tilt to the test tube than this. As the heating commences, a deep blue liquid may move toward the mouth of the upper tube. If this happens, have the student direct the burner flame just in front of the liquid to evaporate it to dryness before it has a chance to escape. Don't be too concerned if some students do not get any water in the vertical tube. All is not lost, and the most important part of the experiment has been salvaged.

2.3 Heating Substances: Part B

Cupric carbonate produces very good results in this experiment. The only problem that is sometimes encountered is that a large lump of the carbonate in the bottom of the test tube does not easily decompose. Have the students spread the material along the tube as shown in the diagram.

2.3 Heating Substances: Part C

We do not usually do this part of the experiment as a class exercise. Interested students generally come in on their own time to do it. The reasons that we do not do this as a class exercise are: (1) each group of students needs two gas jets; (2) the sophistication of the equipment; (3) the flame at the glass jet sometimes flares 20 or 30 cm above the jet itself, making the experiment a little dangerous as a class exercise.

When this experiment is done the glass jet should be ignited before the tube is heated, and when the exercise is complete the bunsen burner is turned off first and then the gas to the glass jet is turned off. All safety precautions should be taken when doing this experiment. Long hair must be tied back and safety glasses worn.

2.3 Heating Substances: Part D

Because the mass of magnesium and magnesium oxide is so small, a reasonably good balance is required to get satisfactory results for this experiment. Sometimes during the heating some of the oxide deposits on the inner lid. Students later remove the lid, and if care is not taken this deposit may be lost. Have the students gently set the lid onto an asbestos mat so that no oxide is lost.

We try to compare data from a large number of trials in each of the above experiments, so we record the data from all of the chemistry classes on a side board. We have found that from a large number of trials the regularity that we are seeking becomes evident.

The answer to question 4, page 82, is of course a statement of the Law of Definite Proportions. It is not necessary to state the Law at this point, but simply to convey the idea that the parts of cobalt chloride were combined in a fixed ratio. Once that idea is established then the next logical step is to test the idea or theory to see if it holds for other

substances – thus the need for the other parts of the experiment.

Gases

A recall of sublimation or dissolving of a solid into a liquid will once again remind the students that certain experiments were explained by assuming the existence of small particles. With this idea in mind, develop with the students a theory or hypothesis about the nature of a gas that will account for something familiar to them: e.g., why does a balloon get larger as you blow into it? Why doesn't the balloon immediately collapse?

The theory should include the ideas that:

1. gases are made up of particles;
2. the *particles* are continually moving;
3. the particles are perfectly elastic.

2.4 Determining the Mass of Gases

This experiment is probably one of the most significant in the unit. The next few experiments all hinge on the results of this one. When students discover that equal volumes of different gases have different masses, each of the following experiments is used to determine *why* equal volumes have different masses, as well as to lend support to the particle nature of gases.

The apparatus used in this experiment can be more easily assembled using a piece of rubber tubing in place of the stopper. This tubing should have an inside diameter of about 6 mm and be cut into lengths of approximately 15 mm. The plastic bags we use are one-quart freezer bags that can be purchased at a grocery store. The gases that we use are:

oxygen,
hydrogen,
tap gas,
propane,
helium,
sulphur hexafluoride.

When we do this experiment we have two lab groups in each class do the propane, helium and sulphur hexafluoride. Their data is recorded on the board for use by all students. In this manner a reasonable amount of data is collected, yet a large amount of gas is not used. All lab groups do the experiment using the first four gases.

The most difficult part of this experiment is to get across to the students the idea that there is a loss of mass due to buoyancy each time the bag is filled with a gas. We have the students do Part A of this experiment, record the apparent mass of air on the board (which of course should be 0.0 g) and then we have a discussion on buoyancy and why the mass appears to be zero. If this idea is not well established here, the chemistry of the experiment may be lost in the resulting confusion.

The data from this experiment is used again in experiment 2.14. Have the students make some special identifying feature in their data table, in outline, in colour or by a special note in the margin.

The answer to question 5 in Part A is especially important and again special note must be made of the theories proposed here. Students will generally propose the ideas given on page 110, or some combination of these two ideas. If other theories are proposed, then all the better. Probably they will be eliminated as more information is collected.

Experiment Values of Density Measured in g/ℓ at 101 kPa and 22° C

Oxygen	*Chlorine*	*Tap Gas*	*Hydrogen*	*Carbon Dioxide*	*Helium*	*Sulphur Hexafluoride*
1.4	2.71	1.78	0.19	1.8	0.3	5.7
1.4	3.11	1.9	0.194	1.66	0.32	6.82
1.33	2.73	1.83	0.224	1.9	0.38	5.11
0.88	2.84	1.35	0.28	1.45	0.15	5.84
1.29	2.76	1.76	0.403	1.82	0.15	6.1
1.19	1.62	1.66	0.29	1.79	0.31	5.9
1.47	*Average*	1.81	0.49	1.69	*Average*	5.0
1.34	2.62	1.72	0.31	1.43	0.26	*Average*
1.61		1.80	0.24	1.29		5.78
1.2		1.92	0.25	1.76		
1.31		1.82	0.15	1.9		
1.30		1.4	0.13	1.78		
1.32		1.82	0.201	1.98		
1.29		1.81	0.38	1.76		
1.25		1.84	0.32	2.91		
1.30		1.89	0.43	0.86		
1.4		1.78	*Average*	1.70		
1.28		1.8	0.28	1.69		
1.47		1.75		1.68		
Average		1.76		1.74		
1.25		1.96		1.83		
		1.76		1.7		
		1.83		*Average*		
		1.8		1.77		
		Average				
		1.84				

Gas Pressure

In the next two experiments you will notice that pressure units are not expressed in kPa but in mm of Hg. It is with special permission that we do this because we feel that the chemistry of the exercises would be lost in the conversion of pressures from mm to kPa.

Students are generally familiar with a mercury barometer and what the length of the mercury column represents. Before proceeding to experiment 2.5 establish the following with students:

1. that if the position of the apparatus (shown on page 91) is changed then the pressure on the confined gas is changed;
2. how to calculate the pressure when the tube is held in the upright, horizontal or inverted position.

2.5 Pressure Change and Gas Volume

It is not our primary intention in experiment 2.5 to arrive at Boyle's Law. Certainly the Law is a consequence of the exercise, but we do it here as an experiment involving gases. In the discussion of the experiment, what will be most important is:

1. if the results can be explained in terms of particles, thus lending support to the particle theory;
2. if the results of the experiment help in the question of different masses for equal volumes of different gases.

Apparatus

Construction of this piece of equipment is not difficult, but it is time consuming.

Use a 40 cm length of 1 mm bore capillary tubing. Close one end of the tube using a propane torch. Use epoxy or contact cement to glue this tube to a 40 cm piece of a meter stick.

Draw a piece of glass tubing into a long slender capillary. Use one end of this as a dropper with a bulb to place a mercury bead into the capillary tube. We have found that a 40 mm bead of mercury in the

tube gives good results for both experiment 2.5 and 2.6. We have used the same set of tubes for these experiments with about 130 students a year for approximately six years.

2.6 Temperature Change and Gas Volume

Once again the primary intent of this experiment is *not* to arrive at Charles' Law but to gather data that will support the particle theory of gases. Certainly the Law does come out of the experiment, and use should be made of it at this time.

The same apparatus is used here as is used in experiment 2.5. A range of temperatures from about 10°C to 80°C produces very good results. We generally have the students get readings for five temperatures within this range. We use Nalgene 250 ml graduated cylinders to hold the water. The tubes *must* be lowered slowly into the hot water or the mercury plug will separate.

Once again this experiment can be explained in terms of small particles, but it does *not* help a great deal with the problem of why equal volumes of different gases have different masses.

2.7 Properties of Gases

To establish the idea that the particles of different gases are different (at least in physical and chemical properties) we do this exercise. This is always done as a demonstration and always with good results. The main point to emphasize here is that the particles of different gases *are in fact different.* This is the first small clue that will lead students to Avogadro's Hypothesis.

Temperature

Students must understand the meaning of temperature to understand experiment 2.8. Further, they must understand how to determine *kinetic energy.*

2.8 Diffusion of Gases

We have recently discovered that this experiment works very well with 60 cm tubes, which eliminate the long waiting time for the reaction to occur. A wooden Q-tip snaps nicely into two parts and presses easily into small corks. The tube *must* be completely dry because of the solubility of the gases used.

The outcome of this experiment is the determination that the particles of the two gases used have different masses. More generally, the particles of different gases have different masses. At this time go back to the theories about different masses for equal volumes of different gases. Experiment 2.7 indicated that the particles of various gases are different, and this experiment indicates that the masses of the particles are different. Students are led to conclude that equal volumes of different gases at the same temperature and pressure contain equal numbers of molecules, their masses being different.

Now that this theory has been established, the logical next step in the scientific method is to test the theory – thus the need for the experiments on reactions between gases.

2.9 Combustion of Hydrogen

This experiment is potentially dangerous and every safety precaution must be taken. We do the experiment in a tank of water as the diagram illustrates, and only with the entire apparatus behind an explosion shield. If an explosion shield is not available, wrap the part of the eudiometer that is not immersed in the water with Scotch tape. This experiment produces good results *if* the oxygen is put in the eudiometer ahead of the hydrogen. We generally do not have a post-lab discussion after this experiment but proceed directly to experiment 2.10 and then to a discussion of how these experiments lend support to the theory that is being tested.

2.10 The Reaction Between Nitric Oxide and Oxygen

The tubes used in this experiment are the same ones used for diffusion of gases. A positive-pressure apparatus for oxygen as well as for the nitric oxide can be assembled for easy dispensing of the two gases. After the two gases are inserted into the reaction tube, the apparatus must be continually agitated until *all* evidence of colour has disappeared.

Students will sometimes find it difficult to understand what the volume of air added after the reaction represents. The post-lab discussion will have to include the chemistry of the reaction so students realize that the air added represents the oxygen used to react with NO. Each group may get two satisfactory reactions before refilling the tube and starting over with a fresh oxygen supply.

Show the Chem-Study film, *Gases and How They Combine,* at this time.

The data in the following tables is actual class data collected from experiments 2.9 and 2.10 respectively.

2.9 Combustion of Hydrogen

Experimental Data Obtained on Six Trials for the Ratio

$$\left[\frac{\textit{volume of hydrogen}}{\textit{volume of oxygen}}\right]$$

Trial	*Ratio*
1	2.07
2	2.0
3	2.03
4	2.1
5	2.0
6	2.05
	Average
	2.04

Within Experimental Error

Ratio $\frac{2}{1}$

2.10 The Reaction Between Nitric Oxide and Oxygen

Results from One Class for the Ratio

$$\left[\frac{\textit{nitric oxide volume}}{\textit{oxygen volume}}\right]$$

Group No.	*Ratio*
1	1.89
2	2.43
3	2.08
4	2.20
5	2.70
6	1.88
7	2.00
8	2.40
9	1.78
10	1.96
11	2.77
12	2.00
	Average
	2.17

Average Ratios from Four Classes

	2.17
	2.12
	1.99
	2.03
	Average
	2.07

2.11 Partial Pressure

The reason for doing this experiment now is two-fold. Firstly, the outcome can be explained in terms of particles of gas and secondly in terms of equal volumes of gas containing the same number of particles.

The apparatus required is somewhat simplified by the use of a side-arm flask. The tube containing the serum cap must be approximately 10 mm inside diameter. To prepare stoppers for this large-diameter glass tubing we put the tube from a cork-borer set into an electric drill. A somewhat difficult and time-consuming task is made simple. Manometers can be purchased mounted on a wooden or metal support. Such devices reduce the chance of a broken manometer and a mercury spill.

The important question in the post lab is number 5, page 119. The results of this exercise are consistent with both the particle theory and Avogadro's Hypothesis.

The idea of names for the particles that we have been talking about may be introduced here.

2.12 Pressure of Liquids

The prime function of this experiment is *not* to introduce vapour pressure or the idea of an equilibrium but to convey that liquids, like gases, may also be made of particles. Since the *results* of the experiment can be explained in terms of particles, the reasonable assumption to make is that liquids may also be made up of these particles.

The apparatus is exactly the same as the one used for partial pressures of gases. A piece of filter paper is placed in the bottom of the flask to absorb the liquid added and provide a larger surface area for evaporation. Sufficient liquid must be added to the flask so that the paper remains damp. The only way that an equilibrium can be established, and thus the maximum vapour pressure, is for liquid to be present in the flask.

To prevent corrosion in the needles they may be washed and then rinsed with a bit of methanol. Pump air through the syringe after the rinsing to evaporate the alcohol.

The Mole

The mole concept is introduced in this manner to be consistent with the remainder of the text, in that many of the topics are introduced in order of their historical development, as opposed to that which may be pedagogically correct. The following experiment illustrates the need for some simple means of

"counting out" large numbers of small particles. Students often wish to *count* the 1000 glass beads. We try to discourage them from doing this and encourage them to have some confidence in the mass method or the volume method of counting out large numbers of particles.

The mole concept leads nicely into atomic and molecular masses, which we introduce at this time. At least one period is spent here manipulating the mole along with atomic and molecular masses.

2.14 The Volume of 1 mol of a Gas at STP

Data from experiment 2.4, which made use of the gas bag assembly, is used in this exercise. Because there generally are students who do not keep a tidy notebook and who may find it difficult to locate the necessary data to do the calculations, we provide our students with an outline to follow and the average class data from experiment 2.4. In this manner every student is dealing with the same data and all will presumably be using it in the same manner. See the outline below.

Sample Outline for Experiment 2.14

Results from the Gas-Bag Experiment (Average of 7 Classes)

Gas	*Pressure*	*Temp.*	*Mass*	*Volume*
tap gas	750 mm	22°C	1.68 g	1000 ml
oxygen	750	22	1.30	1000
hydrogen	750	22	0.28	1000
carbon dioxide	750	22	1.72	1000
helium	750	22	0.26	1000
propane	750	22	1.73	1000

Note: All were determined for 750 mm and 22°C. The volume should be converted to STP for comparative purposes.

1. Convert the 1000 ml to STP.

2. Determine the GMM for each gas.
   C_3H_6
   O_2
   H_2
   CO_2
   He
   C_3H_8

3. Determine the number of moles of each gas used.
   tap gas
   oxygen
   hydrogen
   carbon dioxide
   helium
   propane

4. Determine the volume of 1 mol of each gas.
   tap gas
   oxygen
   hydrogen
   carbon dioxide
   helium
   propane

Unit 2: Sample Test

1. Write formulas for the following compounds.
   (a) carbon disulphide _____
   (b) sulphur dioxide _____
   (c) diphosphorus pentoxide _____
   (d) sulphur hexafluoride _____

2. Show that the mass of a mole (the molecular mass) of CO_2 is 44.0 g and the mass of a mole of SO_2 is 64.1 g.
   _____

3. Calculate the mass of 6.02×10^{23} molecules of carbon monoxide.
   _____

4. How many moles of iron atoms are present in 1.73 g of iron?
   _____

5. What elements do the following symbols represent?
   K _____
   Ca _____
   Co _____
   C _____
   Pb _____

6. Give the name of each of the following compounds.
   HI _____
   BCl_3 _____
   CCl_4 _____
   OCl_2 _____

7. How many moles of atoms are in
   (a) 18 g of Al _____
   (b) 1.83 g of Fe _____
   (c) 56 g of Si _____

8. Suppose chemists had chosen a billion billion (10^{18}) as the number of particles in a mole. What would the molecular mass of oxygen be?
   _____

9. How many atoms are contained in a mole of each of the following?

   H_2SO_4 _____

   $Ca(OH)_2$ _____

   CH_3COOH _____

   $CuSO_4.5H_2O$ _____

10. How many molecules are there in 16.0 g of methane (CH_4)?

    _____

11. How many moles of water molecules are there in 180 g of water?

    _____

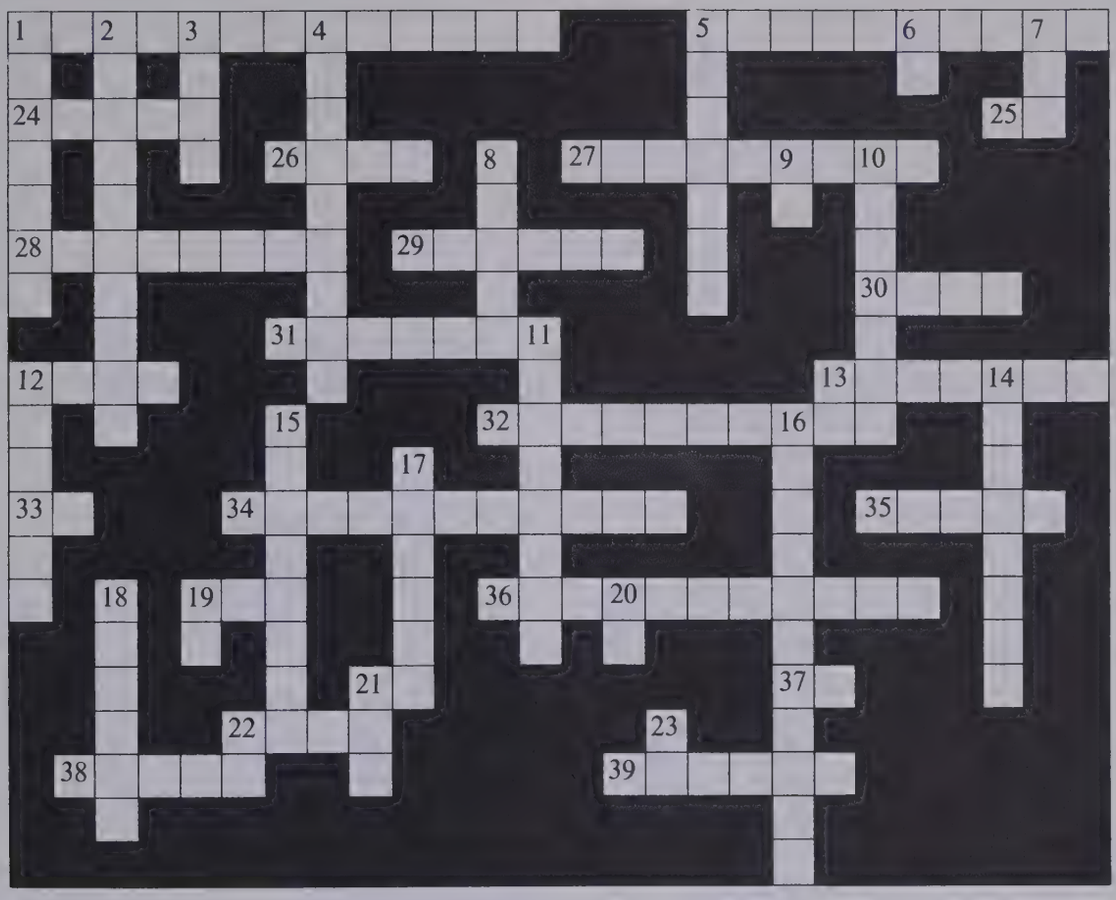
12.

Ammonia	*Hydrogen Chloride*	*Ammonium Chloride*
NH_3	HCl	NH_4Cl
0.17 g	0.365 g	0.535 g

(a) Express the combining proportions in the above chemical reaction in terms of moles.

(b) Express the ratio in the smallest whole numbers.

Unit 2 Crossword Puzzle



Down

1. mass/volume
2. not unlike burning
3. Use a balance to determine this.
4. opposite to decomposition
5. substance taking part in a chemical reaction
6. atomic number is 18
7. used in solder
8. Add an "ar" to the abbreviation for mole.
9. symbol for nickel
10. a kind of fur
11. left side equals right side
12. cowboys and _____ (element)
13. atomic mass is 9
14. a light metal
15. force/area
16. average molecular energy
17. absolute or _____
18. short form for an element
19. 31
20. osmium (symbol)
21. an oil additive
22. raises weather balloons
23. When they made the six-million-dollar man the hardest part was to put his _____ (symbol).

Across

1. brought about by heat or electricity in our experiments
5. constant
24. Boyle, Charles, Avogadro are examples
25. symbol for 12 down
26. smallest part of an element
27. The name of this element is often confused with magnesium.
28. cylindrical container (two words)
29. first name is John
30. _____ substance
31. _____ solutions and pure substances
12. strong metal
13. Something that floats is this.
32. used for combustion experiment
33. 77
34. Yours is 37°C.
35. join to make molecules
19. liquid, solid and _____
36. even throughout
37. symbol for 14 down
38. Robert _____
39. Latin for iron
22. a form of energy

Unit 3
The Atomic Model of Matter

Most students already know all about atoms. Most students do *not* know why they know about atoms. This unit traces the growth of the atomic model from its form as Dalton imagined it to the form in which Bohr, Lewis and others of that period imagined it. If one thing is important in this unit it is the continuation of the idea established in Unit 2 that a scientific model is a growing, changing thing. There is no ultimate model; but each new step is an attempt to improve the previous model. Each model is useful. Each model helps to explain something.

At the end of Unit 2 students should have grasped the idea that there are substances we have called elements that may be composed of small particles called atoms. If these atoms exist, then we have some ideas about them. They may join with other atoms to form compounds. Unit 3 tries to improve the student's concept of these basic particles and to expand his awareness of the scientific process.

3.1 The Black Box

Seeing is believing! Or is it? Not necessarily, as this lab is intended to show. Remind your students that we developed the molecular model without seeing a molecule; that we can believe without seeing if the evidence is there, provided by our other senses.

Each student or pair of students gets a sealed box. A cardboard box about 10 cm × 10 cm × 20 cm does very nicely. Each box should have about three objects in it, such as rubber stoppers, corks, glass tubing, marbles, bolts, nuts, washers, styrofoam spheres, and so on. Try to put objects that have different shapes and are made of different materials in each box.

Do not let the students open the box even when they have finished the lab. Otherwise the idea that we can describe without seeing is destroyed.

Go over the questions that follow the lab before they actually begin. This lets them know some of the things they should be looking for.

Some years ago one of our students took an X-ray of a black box. This is what gave us the idea of including an X-ray on page 139. Note that one object is *very much like* a two-hole rubber stopper: the purpose of this lab is not to guess what the objects are but to try and describe them.

3.2 Static Electricity

Your students need to be reminded of the properties of static charges. If you feel they are not knowledgeable on this topic, then do this experiment. All the materials should be available in the physics lab.

To make an insulated handle for the pie plate, just stretch two elastic bands across the diameter of the plate, at right angles to each other. Pinch the intersection of the elastics between your fingers for the handle.

Following the lab, study the section called "Electricity and Matter," which begins on page 142. Tie matter and electricity together and finish with the Thomson model of the atom. Make sure that your students realize that the Thomson model is an improvement over the Dalton atom model because it helps us explain static electricity.

3.3 Cathode Ray Tubes

Do this demonstration if it is at all possible. The most obvious value of the demonstration is the reinforcement of the idea established earlier – that matter and electricity are somehow connected. It also should bring out the idea of sub-atomic particles, electrons and, by deduction, protons. Information about the size and charge of these particles comes from Table 3.1 on page 150. We see no particular reason for specifying the mass of the electron too accurately. Be sure to finish up by pointing out that these ideas added something to the model but that the model describing the nature of matter will probably never be complete.

Somewhere in here we must point out that Robert Bunsen – yes, *that* Bunsen, – said, "The chemist who is not also a physicist is a fool." Much of the early work in this unit is at least partly physics. You will

do your students a service if you make sure they understand that no science stands alone. Developments in one science usually yield new ideas in other sciences. It is easy to make this point as you go through the "Other Discoveries" section on page 151. This is particularly true of Rutherford's work on radioactive materials.

3.5 The Geiger Counter

If you have a radioactive source and a Geiger counter, show your students the penetrating power of the emanations produced by the source. If they see this it will be easier for them to understand Rutherford's scattering experiment.

Luctor Scientific Company, at 27 Flatte Avenue in Hamilton, Ontario, handles a device made by EME called "The Hypothesis Machine," which makes an excellent pre-lab to the Rutherford lesson. The machine, which sits on the overhead projector, allows you to "fire" small spheres at a hidden target and make predictions about the shape by studying the effect the target has on the spheres.

Be sure you show the students why Rutherford's experiment could not be explained by Thomson's model of the atom, and, therefore, why Rutherford had to modify the Thomson atom somewhat to satisfy the data he had gathered. As far as the calculations on page 156 are concerned, we do not expect students to be able to reproduce them. The three points at the bottom of the page tell it all.

Before you leave the Rutherford model be sure your students realize that it too is incomplete. Talk for a moment about Hertz and Marconi and radio waves. Point out that accelerating electrons are supposed to radiate energy and therefore the planetary electrons should end up in the nucleus as a result of this. Rutherford's model is an improvement over Thomson's atom but it is still not complete.

3.6 Atomic Dimensions

This is a very straightforward exercise in determining the size of the atom. Most students believe that atoms are small. Here is a chance for them to see just how small.

Niels Bohr's contribution to the growth of the atomic model was both simple and complex. To us the essence of what Bohr said hinged on the assumption that there are two sets of rules for electrons to follow. One set applies to electrons that are "outside" atoms and the other set applies to electrons that are "inside" atoms. Once we can accept that, the Bohr modification to the Rutherford atom becomes fairly straightforward.

After you have developed the idea of shells, make every effort to show your students some atomic spectra as outlined in experiment 3.7.

3.7 The Atomic Staircase

Once again you should go to the physics lab for this equipment. You will need a high-voltage source and a couple of gas-discharge tubes. Hydrogen and neon are as good as any. If your school teaches PSSC then you can get the mercury vapour lamps, but take precautions for the UV produced by them. Small hand-spectrometers with an inexpensive diffraction grating are available from several sources – such as Macalaster, at about 50¢ each, or less if purchased in quantity.

3.8 Biological Implications

If you decide to do this one your students will see absorption rather than emission. You can use it to make many points, including why grass is green!

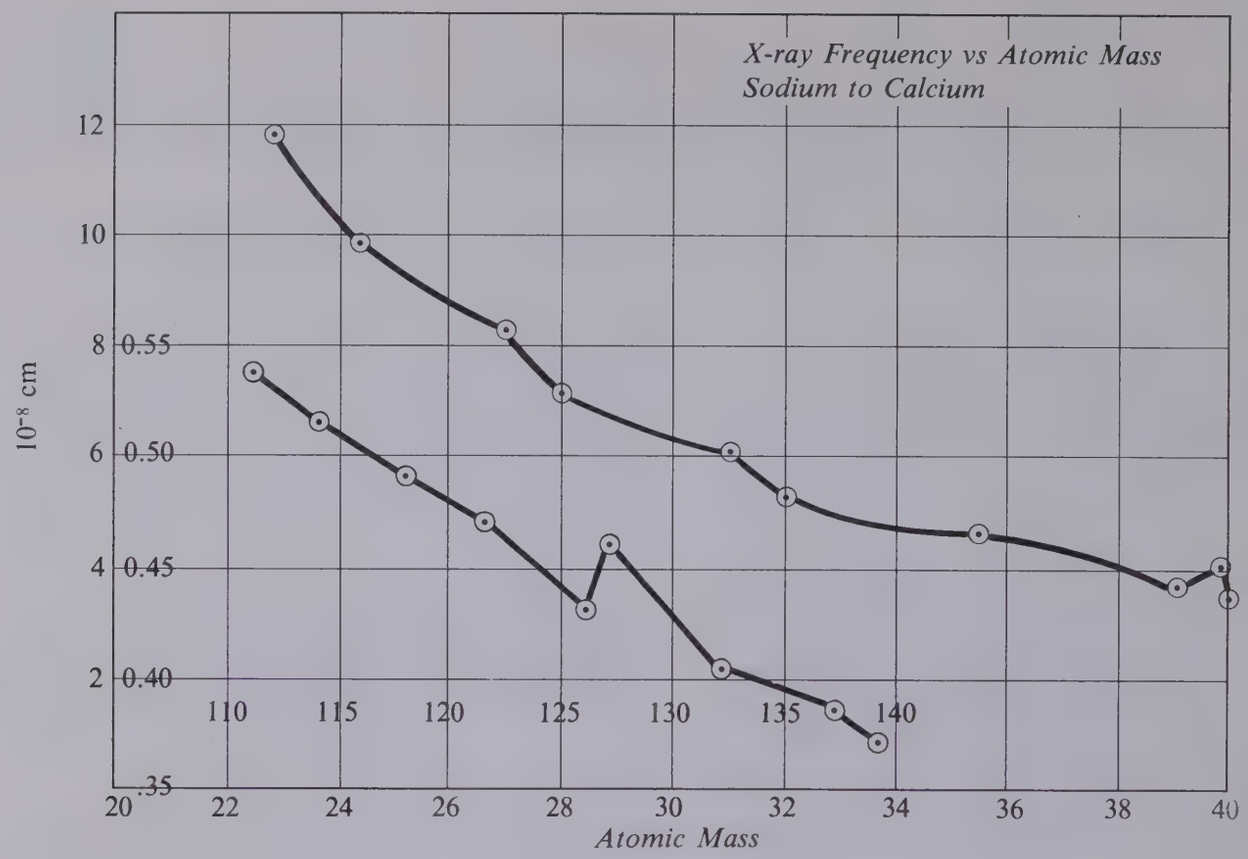
Law of Atomic Number

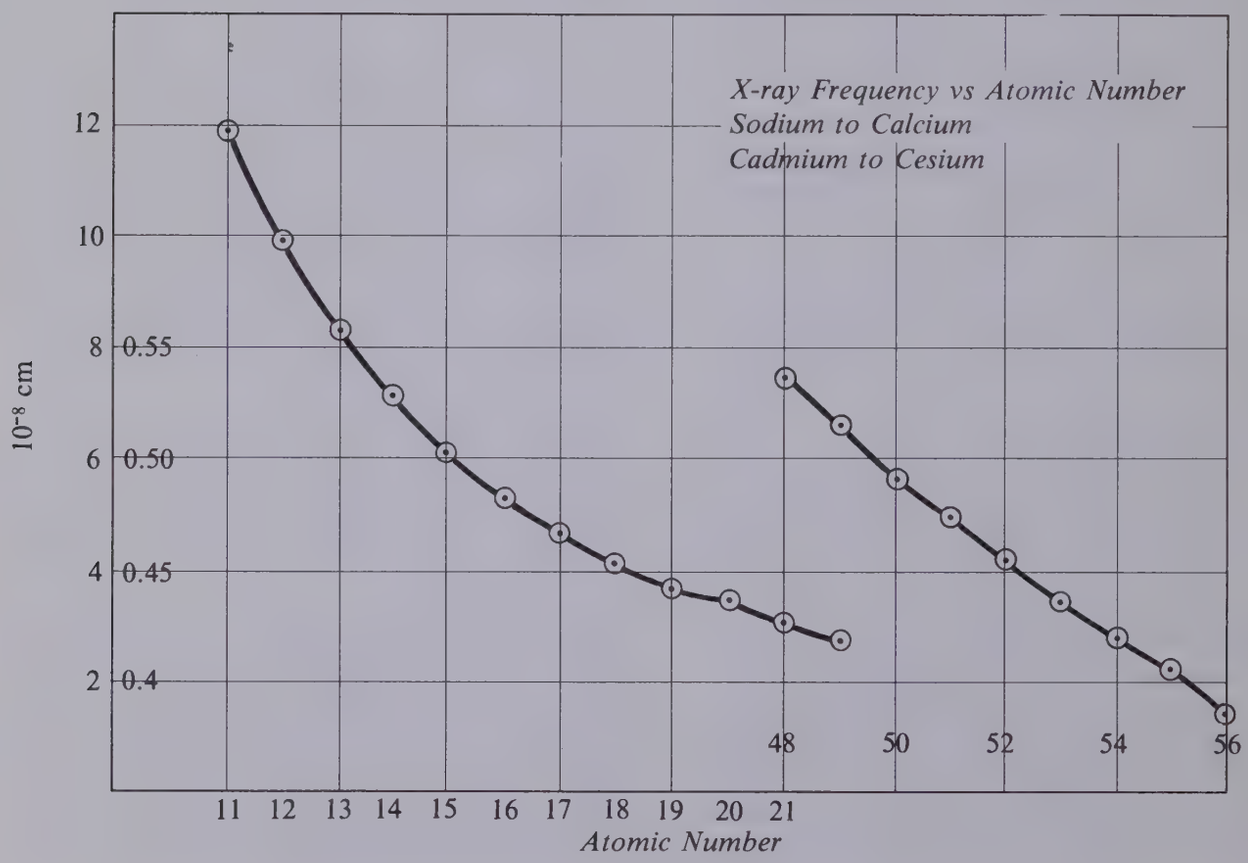
Physics again with Henry Moseley, but most important to the development of the model. The graphs on page 32 may make it easier to understand what Moseley did and why he interpreted the results the way he did.

One way to handle the idea of atomic numbers is to provide your students with the data Moseley had on atomic mass and X-ray emission frequencies and ask them to interpret the results after they have graphed them. We have never had too much luck with this method. We give them the information and then show why Moseley chose the interpretation that he did. It may not be an inquiry approach, but it seems to work here. (If you have not already done so, you should be sure to read *Moseley and the Numbering of the Elements.*)

We have always been happy with the students' discovery of the neutron as it happens in the development on pages 167-169. Your students can suggest a few things about the neutron if you will lead them into it.

Ionization energies provide the data you can use to develop the idea of energy levels. The development on pages 170-174 uses ionization energies to lead students into the concept of maximum occupancies. Don't forget that Bohr had already provided us with shells. Now we can determine the number of electrons in each shell. If a student questions the small dips in the curve, be sure you agree that they


X-ray Frequency vs Atomic Mass
Sodium to Calcium
10⁻⁸ cm
12
10
8
6
4
2
0.55
0.50
0.45
0.40
.35
110
115
120
125
130
135
140
20
22
24
26
28
30
32
34
36
38
40
Atomic Mass


X-ray Frequency vs Atomic Number
Sodium to Calcium
Cadmium to Cesium
10⁻⁸ cm
12
10
8
6
4
2
0.55
0.50
0.45
0.4
48
50
52
54
56
11
12
13
14
15
16
17
18
19
20
21
Atomic Number

are important, but not just yet.

We suggest that you go only so far as calcium in your development. This will permit you to show that the order of filling is not as regular as it could be – there is more here than our simple theory suggests. (Don't forget the small dips in the ionization curve!)

3.9 Conductivity

This is a good opportunity for your students to provide explanations of how the electrons get from one electrode to another. We have provided you with four explanations, but at least three more are possible. Accept the far-fetched ones too!

3.10 Electrolysis

This experiment involves conduction, so the explanations suggested for the last lab can be put to the test here. If you do this as a demonstration with the Hoffman apparatus, then let it run for 30 min or so before the demonstration so the water on the chlorine side gets saturated, otherwise you won't collect much chlorine gas.

Be sure you consider this lab in relation to the explanations provided for the previous one. Does the explanation still work when we consider the new data just uncovered in this new experiment? The Arrhenius theory literally bursts out of the data. Few students will disagree with the idea of ions and ionization. This provides a convenient opening into the concept of valence and electron transfer. Later it can be modified to include sharing.

We do not believe in memory work very often, but we insist that students memorize the valence number for each of the first 20 elements and a similar number of elements beyond the first 20.

3.11 Reactions Involving Solutions

There may be an odour of hydrogen sulphide produced in this lab.

It is important that all the concentrations be close to the specified value, otherwise you may have trouble convincing your class that the precipitates are all the same.

Several things of value come from this experiment.

1. Ions are present even without the electric current. Apparently ionization occurs when the substance is dissolved.
2. Some ions react together and some do not. Those that do not are called spectator ions.
3. We can write an equation showing only the things that react, leaving out the things that do not react.

Note:

We should not be caught recommending that you cheat, but could we tell you that we put sodium sulphide in the bottle labelled potassium sulphide? For reasons we can't explain, it seems to work better that way!

3.12 The Electrolysis of Aqueous Copper II Chloride

Be sure that your students make their predictions before they do the lab. As the electrolysis proceeds a definite copper colour will form on the one electrode and a definite odour of chlorine will be produced.

Types of Bonding

The section on the covalent bond, which starts on page 191, should complete the section on bonding. Please note that we do not talk about architecture or molecular shape. We feel that is best left to the next year of chemistry. Because we don't cover shape we see no need to get into the subject of sub-shells. We feel that sub-shells are very useful to explain molecular shape and the polarity of molecules, but that this is too advanced for a beginning course in chemistry.

If you want to, you can easily bring in the idea of the polar covalent bond by discussing the idea of unequal sharing introduced on page 192. If you feel strongly about it, you can also move into molecular architecture from the section on "Lewis Diagrams," on page 193.

In any case it is probably wise to include the section showing the shared pair as a line, as on page 194. This is the notation used in the organic section.

Problems

When you do the problem section on page 195, silicon carbide will give you an entry into a short discussion of covalent solids or macromolecules. It is probably worth a few moments of your time to mention diamonds and the strength of the covalent bond.

Isotopes and Radioactivity

The sections on "Isotopes" and "Radioactivity" are intentionally short. It is our feeling that we should expose our students to these two ideas, but since they are not fundamental to the ideas we are trying to develop, we do not go too deeply into them. We feel it is enough to let students know that there are other things to learn about in this area.

The following test will give you some idea of the depth of treatment we give to some of these topics.

Unit 3: Sample Test

1. Draw the Lewis Dot Diagram for
   (a) a silver atom
   (b) a fluorine ion
   (c) a neon molecule
   (d) a carbon disulphide molecule
   (e) a magnesium ion

2. Draw the structural formula for the following substances, using a line to represent the shared pairs.
   (a) water
   (b) ammonia NH_3
   (c) acetylene C_2H_2
   (d) propane C_3H_8
   (e) carbon dioxide

3. Write the balanced chemical equation for each of the following chemical reactions.
   (a) magnesium combining with sulphur
   (b) the synthesis of carbon IV oxide
   (c) the ionization of sodium bromide
   (d) the decomposition of lead IV oxide into lead II oxide and oxygen gas
   (e) the reaction between sodium sulphide and copper fluoride

4. Balance the following equation.
   $Na_2CO_3 + H_3PO_4 \longrightarrow Na_3PO_4 + H_2O + CO_2$

5. Draw a diagram to represent
   (a) a sodium atom
   (b) a calcium atom
   (c) an inert gas atom
   (d) a lithium ion
   (e) a fluorine ion

6. Complete the following table by writing in each space both the name and the formula for the substances formed. Note the example given in the first row.

Chlorine	*Sulphur*	*Nitrogen*	*Carbon*	
NaCl Sodium Chloride				Sodium
				Magnesium
				Aluminum
				Hydrogen

Unit 4
Descriptive Chemistry and Stoichiometry

By the beginning of this unit (if the text has been followed) students will have been exposed to reasons for our belief in the existence of atoms and molecules, and to some of the basic concepts of atomic structure and bonding. In addition, they will know valence numbers for many of the elements and how to use these for formulas of binary compounds. The intent of Unit 4 is to introduce the students to changes in matter (chemical reactions) and to the writing and use of equations related to these reactions.

In the process of studying these reactions several common elements or compounds will also be examined. In addition, Unit 4 introduces acids and bases as well as reactions of acids and bases.

We have found that as we teach this unit we get the best results if we treat it as a study of reactions rather than treating each of the different topics of study as an isolated unit. The study of reactions will tie all the topics in the unit together.

This unit is quite descriptive in places, but, like the rest of the book, the actual textual material has been reduced to a minimum. So the students must observe well and record accurately in the 22 experiments they will be expected to perform.

4.1 The Effect of Manganese Dioxide on Hydrogen Peroxide

With a good balance this experiment will produce excellent results. We have found that Part A of the experiment works best with powdered manganese dioxide, while Part B gives best results with granular manganese dioxide. The powdered MnO_2 provides a larger surface area for reaction to occur, but if it is used in Part B the fine particles clog the filter while some may even pass through the filter.

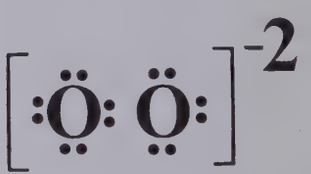
Remind the students that for Part B of this experiment it is *not* important to have exactly 5 g of MnO_2. What is important is that they know exactly *what* the mass of the amount they select is.

Manganese dioxide is a messy material to work with. Students should be told this and asked to avoid spills on the counter, floor or themselves. When washing containers in which they have had MnO_2, the mixture should be poured into the drain and flushed down with water.

We have found that the post lab to this exercise is a convenient time to introduce the writing of formulas of peroxides. The question that is most often asked by students here is about the type of bonding in peroxides.

Peroxides

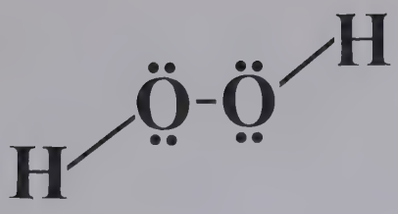
Oxygen reacts with some of the group IA and IIA metals to form peroxides in which oxygen has an oxidation state of –1. When sodium or potassium or barium are heated under pressure with an ample supply of oxygen gas, the main products are Na_2O_2, K_2O_2 and BaO_2 respectively. The structure of the peroxide ion is:



If solid peroxides react with acid solutions, hydrogen peroxide is produced.

$Na_2O_2 + 2H_2O^+ \longrightarrow 2Na^+ + 2H_2O + H_2O_2$

Hydrogen peroxide is a clear, colourless, syrupy liquid. It melts at –1.7°C and boils at 155.5°C. It decomposes spontaneously to yield oxygen and water. The decomposition rate is accelerated considerably by some alkalis, heavy metals and dust particles. Hydrogen peroxide is most often used in a 3 per cent solution or sometimes as a 6 per cent solution. The bonding in hydrogen peroxide is:



The chart below gives actual class data obtained in this experiment.

Group	*Paper + MnO_2 Before*	*Paper + MnO_2 After*	*Change in Mass*
1	1.68	1.67	0.01
2	1.57	1.58	0.01
3	1.53	1.51	0.02
4	3.42	3.42	0.00
5	1.01	0.99	0.02
6	1.16	1.16	0.00
7	1.74	1.74	0.00
8	1.23	1.34	0.11
9	2.27	2.26	0.01
10	1.20	1.20	0.00
11	0.71	0.97	0.26
12	1.23	1.26	0.03
13	2.68	2.71	0.03
14	1.29	1.64	0.35

4.2 Properties of Oxygen

This experiment (not including the demonstration) can only be done in a 45 min period if the students are well briefed on both the preparation and the use of the oxygen that they collect. We find it convenient to set up experiment 4.3 in part of a period and then use the remaining time for a pre-lab of experiment 4.2. Time should also be available in this period for the students to prepare a data sheet, certainly a necessary part of the experiment.

Preparation

We have found that if we prepare the potassium chlorate-manganese dioxide mixture ahead of time a few precious minutes are saved for the students. It is important that the students take the 3 cm of the mixture as specified in order to collect the four bottles of gas. The delivery tube must be clean to allow gas to escape freely. The stoppers on the delivery tubes should be changed occasionally, because as they become hard and brittle small bits may fall into the mixture of potassium chlorate and manganese dioxide, causing a rather violent and potentially dangerous reaction. Draw the students attention to the warnings given on page 214.

Using the Oxygen

Distilled water should be used in procedure 5, since the resulting solutions will only be slightly acidic or basic. Only small lumps of sulphur should be used in procedure 8 and the flame should be extinguished immediately after removing the spoon from the gas jar to prevent excessive amounts of sulphur dioxide from entering the room.

To prevent damage to the desk in procedure 9, an asbestos square should be placed on the counter and the jar with burner placed on this square. The burning magnesium must be lowered into the jar of gas *without* touching the glass or the excessive heat will break the bottle.

In the demonstration for procedure 10 we generally use not only sodium as indicated but also potassium and phosphorus. By using these three elements in addition to those used by the students, data for three metals and three non-metals will have been collected.

4.3 The Proportion by Volume of Oxygen in the Air

The steel wool must be washed thoroughly with water *after* it has been rinsed in the carbon tetrachloride. If this is not done the CCl_4 remaining on the steel wool will evaporate, displacing some of the water that will have entered to replace the oxygen used.

Sample Data from 4.3 Part A

This is actual class data collected from five demonstrations to five different classes.

Total Volume of Air	*Volume of Water Entering Cylinder*	*Percentage of Oxygen*
323 ml	67.8 ml	21 %
329	73.4	22.3
316	72.7	23
309	67.2	21.7

We find this experiment to be useful here because it recalls the ideas of molar volume and gram molecular mass established in Unit 2. If students collect all of the *gas* produced by the $KClO_3$ (this means they must decompose *all* of the $KClO_3$), very good results are obtained. The data given below is

actual class data collected from this experiment. If students use approximately 5 g of $KClO_3$ they will collect about 1200 ml of gas. This means that they must start with at least five standard-size gas bottles. If these bottles are not filled with water and ready for use before the heating begins, then too much gas will be lost to the air.

4.3 Part B

Volume of Oxygen from 1 mol of $KClO_3$ (room conditions)

	35 920 ml
	21 095
	34 348
	30 992
	36 199
	26 013
	34 000
	29 119
	33 967
Average	32 406 ml

4.4 Properties of Acids
4.5 Properties of Bases

We have found that the best use of our time is made if experiments 4.4 and 4.5 are done on successive days, followed by a discussion of the outcome of the two experiments. Experiment 4.4 is intended to demonstrate to the students the following properties of acids:

1. They have a sour taste.
2. They affect certain indicators, i.e., litmus, bromthymol and phenolphthalein.
3. They react with some metals to produce hydrogen.
4. They react with carbonates liberating CO_2.

Experiment 4.5 is intended to demonstrate to the students the following properties of bases:

1. They taste bitter.
2. They feel slippery or oily.
3. They affect indicators.

With properties of acids and bases established this is a convenient time to establish names and formulas for both acids and bases.

4.7 Electrical Conductivity of Acids and Bases

Since the students have already had the opportunity to do a conductivity experiment in Unit 3, we have found that time and materials can be saved if this exercise is demonstrated by you. The intent of the experiment is to relate degree of ionization to acid strength. In Unit 3 it was established that solutions conduct electricity only if the solute dissociated into ions. It therefore follows that the brighter the light, the higher the concentration of ions. An acid, for example, that has a high concentration of ions in an aqueous solution is a strong acid. Similarly a base with a high concentration of ions in an aqueous solution is a strong base. Conversely, weak acids or bases have low concentrations of ions in an aqueous solution.

This is also a convenient place to introduce the students to the idea of hydrated ions.

4.8 Neutralization Reaction between an Acid and a Base

This exercise is intended to illustrate that a product of a neutralization reaction is a salt. In order to identify the residue as NaCl we have asked the students to taste a small portion of the residue. To ensure that the residue contains little or no NaOH it will be important that all of the base is neutralized. If, during the heating, a pink rim forms in the evaporating dish, have the students add *one* drop more of the acid and stir it and the residue forming the rim into the solution. With a bit of care this can easily be done while heating continues.

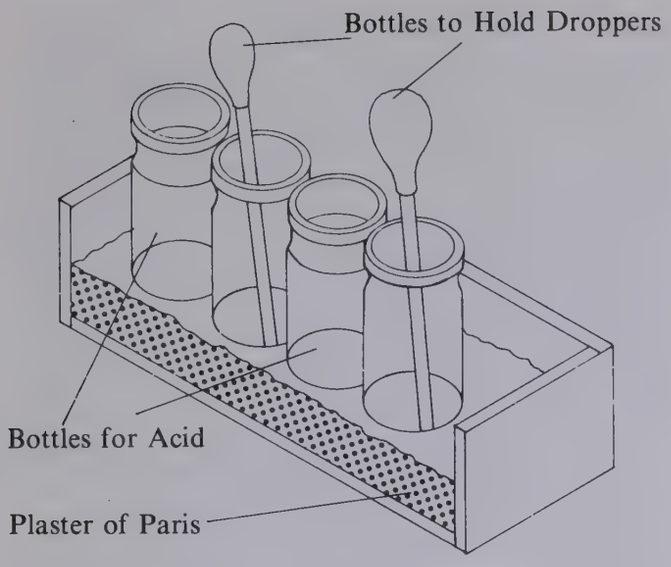
4.9 Preparing a Molar Solution

This is an excellent exercise for providing students with one of their few opportunities to prepare a solution. It also provides them with a chance to use a volumetric flask and to handle potentially dangerous chemicals. For the teacher it presents several problems:

1. how to dispense the NaOH;
2. how to dispense the HCl;
3. where to store the solutions.

To dispense the NaOH we place close to the correct number of pellets in small vials with screw caps. As the student uses the amount he needs he returns it to the teacher for re-filling. This procedure tends to eliminate crowds around one bottle, spilled pellets and a wet bottle of NaOH pellets.

To dispense the acid we have prepared wooden trays filled with plaster of Paris, into which we have embedded a reagent bottle and a dropper bottle. The reagent bottle is filled with acid while the dropper bottle is used only to hold the dropper. Several such boxes placed about the room eliminate the need for students to carry any large quantity of acid from one place to another. It also eliminates the need to place a dropper, wet with concentrated HCl, on the counter.

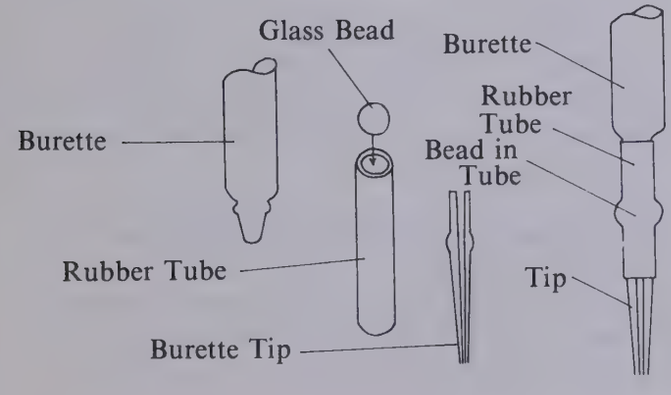


We ask each student to bring two small jars or bottles from home to store their solution in. A request for these several days ahead of time generally provides adequate containers at experiment time. The containers should have tight-fitting lids. Baby-food jars are excellent for this purpose.

For a reasonable understanding of the experiments and work that follow this exercise, the students must at this point develop an adequate grasp of the concept of molarity and molar solutions.

4.10 Acid-Base Titration

This is an excellent exercise for introducing students to the use, handling and care of burettes and pipettes as well as to the calculations of concentrations from titration data. We have found that glass burettes with ground-glass stopcocks and grease present innumerable problems at the early stages of titrations. We have had good results with glass burettes and rubber-tube glass-bead controls and excellent results with plastic burettes and teflon stop-cocks.



4.12 Acidity of Some Common Substances

Coloured solutions may present problems in this experiment, since the colour can interfere with the indicator selected. We have found that confusion and problems can be eliminated by using colourless liquids such as vinegar, Sprite, ginger ale or pickle juice.

The acid or base that is used for neutralization of these common substances should be fairly dilute. We have found that an acid or base that is 0.1 M gives good results.

4.13 Back Titrations

The actual doing of this experiment is not difficult, but difficulty often arises with the calculations that follow. The students already have an understanding of neutralization and of the idea that in neutralization the number of moles of hydronium ion must equal the number of moles of hydroxyl ion provided from the base. In this back titration it may help to convey this idea in the form of an equation.

moles of hydronium ion from acid = moles of hydroxyl ion from base solution + moles of hydroxyl ion from pellet

or

moles of hydronium ion from acid − moles of hydroxyl ion from the base solution = moles of hydroxyl ion from pellet.

Since the volumes and molarites of both acid and base are known, then the moles of hydronium ion and moles of hydroxyl ion may be calculated. Subtraction yields moles of hydroxyl ion provided by the pellet. Students must have a reasonable understanding of back titrations in order to make any sense out of the antacid analysis.

4.14 Investigation of Several Commercial Antacid Preparations

This exercise may be done as a mini-project. Provide the student with eight to ten antacid preparations to examine. Allow three or four periods to carry out the examination and then have students present a report.

In this case the students may be encouraged to carry out more investigations than those suggested:

e.g., solubility, insoluble residue present or temperature change upon reaction.

The outline that follows is the sort of thing we give to our students at the start of this exercise.

Consumer Chemistry

In simplest terms the purpose of this assignment is to investigate several commercial antacid preparations in whatever ways seem appropriate to you. All groups will follow the same basic procedure to determine the antacid capacity of each preparation. Additions to the procedure are welcome. You may wish, for example, to include in your study reaction rate, tablet mass, insoluble residues or some other parameter.

Eight preparations are available, seven of them in tablet form. Brand X has been developed by the chemistry department and is in the early stages of testing. You are allowed a maximum of three tablets from any one bottle. Note that Brand X is not in tablet form.

You may prepare a stomach from a large test tube by adding 10.0 ml of the stomach acid provided. Note the molarity. Add one antacid tablet and allow the reaction to subside. Add indicator and back titrate the excess acid with a standard base. Repeat each brand twice more. *You are expected to attempt at least five different brands.*

Price information, if you need it, is as follows:

Cost per 100 Tablets:

Rolaids	$1.69
Tums	1.09
Creamalin	1.95
DiGel	2.10
Bisodol	1.09
Phillips	.99
I.D.A.	.69

Brand X is available at $1.39 per pound.
Your teacher will supply you with any other information you may need.

4.15 The Reaction of Some Metals with Dilute Acids

The students should be instructed to handle the calcium with a pair of tongs. If the calcium has been stored for some time it will be heavily oxidized. If this is the case have the students expose some of the bare metal by filing some of the oxide away. It may also be a good idea to run the magnesium ribbon between some emery paper or fine steel wool to remove the oxide from it as well.

4.16 Investigating Equations

Because students are using such a small amount of magnesium in this experiment, we have found that much better results are obtained by providing them with the mass of some greater length (100 cm) and having them calculate the mass of the strip they use. As in the previous experiment the oxide should be removed from the magnesium ribbon. Students generally find the calculations to this experiment difficult and will have to be helped through them.

The chart below provides the vapour pressure of water at various temperatures.

We are using this experiment to lead into the idea of mass-volume relationships from equations. To answer question 6 we are assuming the student knows that the molar volume is 22.4 ℓ at STP.

Vapour Pressure of Water mm Hg

°C Temperature	*Vapour Pressure*	*°C Temperature*	*Vapour Pressure*
10	9.209	21	18.650
11	9.844	22	19.827
12	10.518	23	21.068
13	11.231	24	22.377
14	11.987	25	23.756
15	12.788	26	25.209
16	13.634	27	26.736
17	14.530	28	28.349
18	15.477	29	30.043
19	16.477	30	31.824
20	17.535		

4.17 Preparation and Properties of Hydrogen

Because of the explosive nature of hydrogen gas, this experiment must be done without open flame near the generator. To prevent the gas from escaping via the thistle tube, the end of the thistle tube should be near enough to the bottom of the flask to be covered with solution.

The students' attention should be drawn to procedure 8 so that they will observe closely when performing this part of the experiment. The splint does not burn in the gas, but the gas burns quietly at the mouth of the jar.

4.18 Reaction between Hydrogen and Copper II Oxide

This is an exciting experiment to perform. However, we feel it is too dangerous to perform as a class experiment. The greatest danger arises when the

bunsen burner is ignited before relatively pure hydrogen is being liberated.

This experiment is done here for two reasons. First of all, it involves the use of hydrogen gas, which has been the topic of discussion in the last two experiments. Secondly, it is one of the few reduction experiments that can easily be performed in a high school lab. If you are willing to forgo the hydrogen part of this experiment and wish to demonstrate the reduction only, then experiment 2.3, Part C, may be done in place of this experiment. The procedure is similar except tap gas replaces the hydrogen and the excess gas is burned off at the end of the test tube, as illustrated in the diagram. Because tap gas does not ignite as explosively as hydrogen the experiment is considerably safer to perform.

When dry hydrogen is passed over the copper II oxide, the oxygen is removed from the oxide to produce water and leaves copper behind in the tube.

$H_2 + CuO \longrightarrow H_2O + Cu$

The post-lab of this experiment leads into a discussion of oxidation and reduction. When we teach this part of the course our discussion of this topic is only brief and oxidation and reduction are defined in terms of a loss or gain of electrons respectively.

4.20 Preparation, Collection and Properties of Carbon Dioxide

This exercise works well as a class experiment. The apparatus is identical to that used for the preparation of hydrogen, so the thistle-tube assembly will already be prepared. This experiment is intended to show some of the properties of carbon dioxide. The fact that it can be collected by the downward displacement of water illustrates its limited solubility in water.

The gas is colourless and odourless with a density greater than air, as is illustrated by the extinguishing of the candle flame as the CO_2 is poured over the burning candle. The gas, of course, does not burn nor does it support burning of an ordinary wood splint.

Carbon dioxide is identified by passing it into limewater, where a white precipitate is produced, giving the limewater a milky appearance.

$CO_2 + Ca(OH)_2 \longrightarrow CaCO_3 \downarrow + H_2O$

The $CaCO_3$ is the precipitate.

If CO_2 is passed through the mixture of water and $CaCO_3$ a colourless solution results by the formation of a soluble bicarbonate.

$CO_2 + H_2O \longrightarrow H_2CO_3$

$H_2CO_3 + CaCO_3 \longrightarrow Ca(HCO_3)_2$ (soluble)

4.22 Dry Ice

This experiment is probably best demonstrated. The procedures are intended to show the class that dry ice displays the properties of carbon dioxide, hence their answer to question 2.

Procedure 1 illustrates that the solution of dry ice (CO_2) and distilled water is acidic.

$CO_2 + H_2O \longrightarrow H_2CO_3$

Procedure 2 illustrates the formation of a white precipitate in limewater, hence a milky appearance.

Procedure 3 illustrates the sublimation of dry ice and the fact that it becomes gaseous at room temperature.

Procedure 4 illustrates that the gas formed by dry ice neither burns nor does it support burning.

Procedure 5 illustrates the carbonating effect of dry ice on water.

Lab Test

The following is an outline of a lab test that we do following experiment 4.20. The five substances that we use are:

1. sodium carbonate
2. sodium chloride
3. sugar
4. powdered magnesium and iron II oxide
5. powdered calcium carbonate and iron II oxide

Unit 4: Sample Lab Test

Confidential
This is a laboratory test. You may not discuss anything with anyone except your teacher and your lab partner.

You are a forensic scientist working in the labs of Inter-Pill, the international anti-drug squad. An informer has told your department that the terrible drug callumine carbonate is being smuggled into Brighton by a sinister group of criminals intent on turning the entire village population into addicts. The informer has also indicated that the shipment is being carried in on the 08:15 bus from Toronto.

Action
The bus has been stopped and everyone on it has been searched. Five of the passengers were carrying suspicious-looking substances and have been detained. As the lab chemists, you must test these five substances and decide which, if any, of them is callumine carbonate. Each sample is labelled with the name of the person carrying it. Carry out your tests and tell the chief who he should detain for further questioning. Remember that your evidence may have to stand up in court. Hand in your report to the chief on the form provided.

Names of Investigators	
Description of Tests Carried Out	
Results of Tests Carried Out	
Recommendation to the Chief	

Unit 4: Sample Test 1

1. If 8.0 g of NaOH were dissolved in 200 ml of solution the molarity would be
   (a) 0.1 M
   (b) 0.2 M
   (c) 1 M
   (d) 2 M
   (e) 0.04 M
2. The gram molecular mass of nitric acid is
   (a) 65 g
   (b) 64 g
   (c) 63 g
   (d) 79 g
   (e) 47 g
3. The number of grams of $CaCl_2$ required to make 500 ml of 4 M solution would be
   (a) 111 g
   (b) 222 g
   (c) 444 g
   (d) 55.5 g
   (e) over 1000 g
4. A 10 ml sample of an unknown acid is titrated with 0.1 M base. The average of three titrations shows that 17.3 ml of base are required for each titration. The molarity of the unknown acid is
   (a) 0.1 M
   (b) 0.173 M
   (c) 173 M
   (d) 1.73 M
   (e) 0.0173 M
5. If the formula for potassium permanganate is $KMnO_4$ write the formula for hypo manganous acid. _____
6. If telluric acid has the formula H_2TeO_4 write the formula for pertelluric acid.
   _____
7. Write the balanced chemical equation for the following reactions. Use the back of this page.
   (a) nitric acid and potassium hydroxide
   (b) phosphorous acid and barium hydroxide
   (c) hypochlorous acid and ammonium hydroxide
8. Write the net ionic equation for the reaction between hydrosulphuric acid and aluminum hydroxide. _____

Unit 4: Sample Test 2

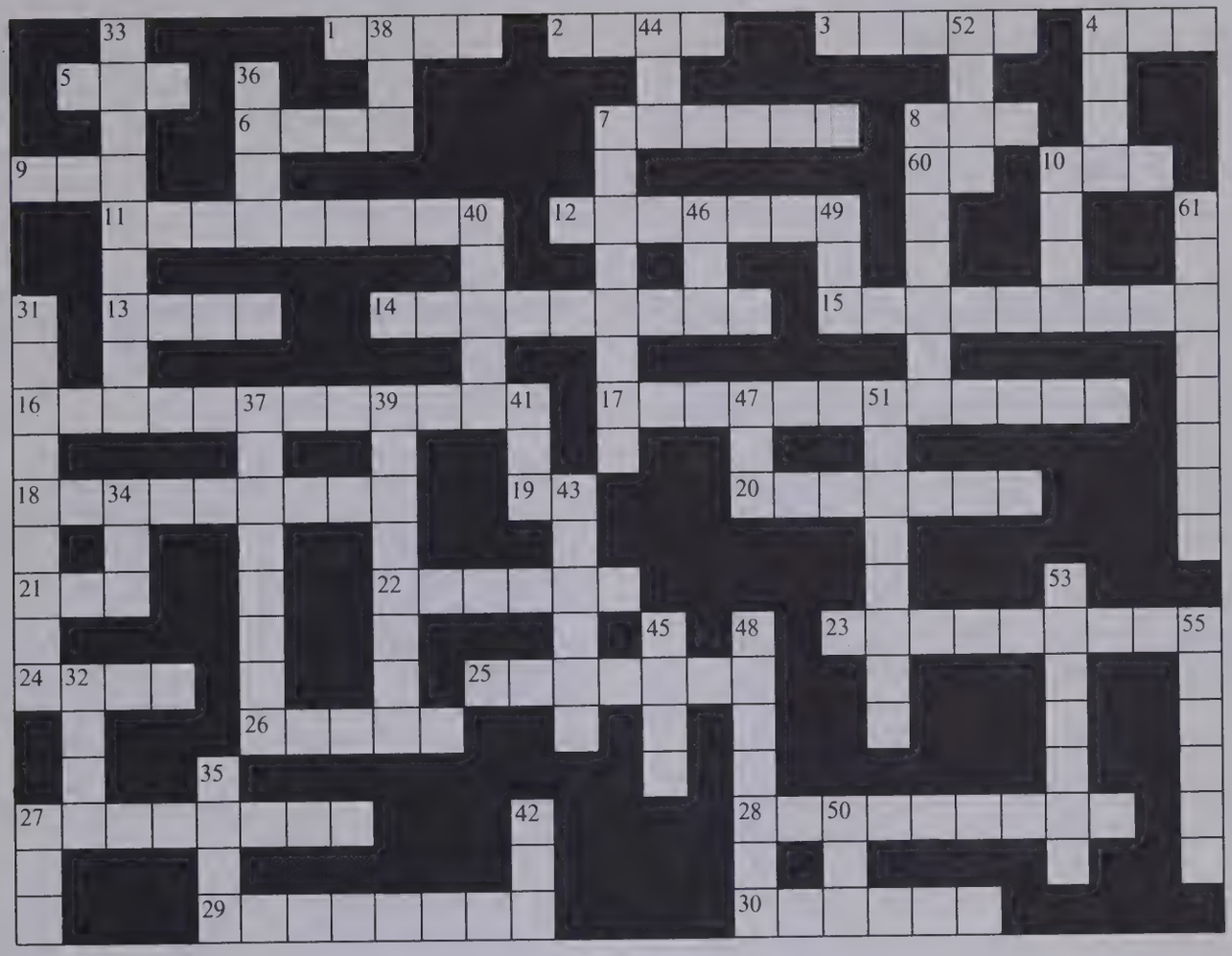
1. Write the balanced chemical equation for the heating of potassium chlorate using manganese dioxide as the catalyst. Explain the function of the manganese dioxide.
2. Write the name and formula for the substance produced when carbon is burned in air and the product is dissolved in water.
3. Grapefruit, vinegar and green apples all have a tart taste. What other properties might they have in common?
4. In a titration one would use a burette, a pipette and an indicator. What is a titration? What is the function of the burette, the pipette and the indicator?
5. Hydrogen reacts with hot copper oxide to produce metallic copper and water.
   (a) Write the balanced equation for the reaction.
   (b) Write the oxidation half-reaction.
   (c) Write the reduction half-reaction.
6. Write a chemical equation for the preparation of carbon dioxide from nitric acid and lithium carbonate.
7. Name the following substances.

   HBr _____ $Ba(OH)_2$ _____
   HNO_2 _____ $FeCl_3$ _____
   $CaSO_4$ _____ $HClO$ _____
   H_3PO_5 _____ $Al(BrO_4)_3$ _____
   Na_2SO_3 _____ $CsOH$ _____

8. Write the formula of

   sodium phosphate _____ hydrosulphuric acid _____
   ammonium hydroxide _____ iodic acid _____
   hypochlorous acid _____ sodium bicarbonate _____
   potassium acetate _____ strontium nitrate _____
   lithium perchlorate _____ barium hyposulphite _____

Unit 4 Crossword Puzzle



Across

1. indicates a low degree of ionization
2. prefix that means "less than"
3. positive electrode
4. beverage that chemists make in beakers
5. fluid of indefinite volume
6. produces H^+ ions
7. added to the front of a name
8. turns blue in a base (abbrev.)
9. aspirin tablets
10. common conditions for a gas
11. used to test for CO_2
12. not mineral
13. used to clean glassware
14. a base
15. experiment to determine concentration
16. one collects a gas by downward _____
17. when one atom replaces another in a molecule
18. combination with oxygen
19. niobium
20. neither acidic nor basic
21. charged particle
22. highly ionized
23. contains OH group
24. 6×10^{23} particles
25. chemists' shorthand
26. _____ zinc
27. very light gas
28. salt of carbonic acid
29. must be balanced
30. common indicator
60. we

Down

32. feel of a base
31. H_3O^+ ion
27. much molecular motion
33. using a catalyst
34. atoms that gain or lose electrons
35. chemist's system of naming is like a _____
36. opposite of acid
37. 89
38. moles acid = moles base is the _____ point
39. curved surface of a liquid
40. radioactive inert gas
41. found in stannous fluoride
42. isn't this _____
43. only two elements
7. more oxygen than the oxide
44. prefix meaning more
45. bunsen burners need _____
46. non-chemical connective
47. used to record observations
48. complex ion
49. biologist's favourite
50. biologist's second favourite
51. alters the rate of reaction
8. measures liquid volume
52. Lewis loved these
10. + ion of a base and - ion from an acid
53. salt of oxalic acid
4. taste of an acid
61. read the bottom of this
55. gained or lost in a chemical reaction

Bonus:
You should be able to circle the symbols for many elements in the completed crossword.

This crossword consists essentially of new terms from Unit 4.

Unit 5 Man's Greatest Chemical Achievement

The Periodic Table

There are those who feel that the development of the periodic table was man's greatest achievement. There will certainly be those who disagree, but it is true that the periodic table of the elements is a masterpiece of organization and design. Students who are familiar with the layout of the table are certainly in an advantageous position to continue their study of chemistry.

We feel that the development of the periodic table carried out in Unit 5 is true to our "code" in that it follows the inquiry approach. The data that the student gathers relates to two important families in the periodic table: the halogens and the active metals. It follows then that in addition to learning about the periodic table the student becomes familiar with the properties of these two groups of elements.

As you go through this unit keep in mind the properties of the elements and the organization of the table.

5.1 Reaction of Metals with Water

Part A is very straightforward.

Part B

You should cut the three metals yourself in advance of the experiment. Cut cubes of each metal about 3 mm to a side. These cubes should then be placed in a petri dish with about 5 mm of kerosene in it. Each student should be given a teasing pin from the biology lab to help manipulate the samples. It will probably speed up the lab if you demonstrate the appearance of the freshly cleaned lithium, sodium and potassium at the start of the period. This also eliminates the need for the student to return the samples to the petri dish after the examination.

The kerosene should be dried off the samples before reacting them with the water. A simple way to do this is to place the sample in the fold of a paper towel. Make sure your students realize that they must not touch the samples with their fingers.

If you have not already done so you should make sure that your students know how to "smell" a gas without actually inhaling any of it into their lungs. There is a description of a safe way to do this on page 397.

5.2a To Prepare and Study Chlorine

If you do not have proper ventilation equipment then you should not do this experiment. Be sure you tell your class that chlorine is a very dangerous gas and that at one time it was used as a war gas. If you have adequate ventilation the lab is quite safe, as not much gas is produced with the amounts specified. Be sure you have students observe carefully the effect on litmus – there is a surprise in store for them!

5.2b Demonstration of Additional Properties of Chlorine

If you can you should use a bottle of compressed gas as your source for this demonstration. We have found that a lecture bottle is large enough to supply our needs for a couple of years. We once tried a lecture sphere – it kept us in chlorine for about six years. The valve corroded before the bottle was empty and it gave us a nasty scare as the chlorine escaped into the fume hood. Chlorine gas can be neutralized by a solution of photographer's hypo or sodium thiosulphate.

It's fun to test the bleaching action of chlorine on other substances. We always try to find a dandelion blossom or some other flower. You can also show the need for water by using some dry samples.

The following equations represent the mechanism of the bleaching process.

$Cl_2 + H_2O \longrightarrow HCl + HClO$

$HClO \longrightarrow HCl + [O]$ (this is called atomic oxygen)

[O] + coloured molecule $\longrightarrow$ colourless molecule

One of the reasons that chlorine bleaches are hard on some materials is because of the acid produced. In commercial bleaches the active substance is frequently NaClO. One often sees this written as NaOCl.

5.3a To Prepare and Study Bromine

The same precautions with respect to fumes apply to this experiment. In addition to the bromine produced, the student is asked to use carbon tetrachloride in this experiment. Be sure you give the necessary precautions for handling the latter.

The equations involved in this lab are similar to those for the chlorine lab. Make sure that your students see the very close relationship between the experiments.

In the demonstration of 5.3b you can add the same sort of thing as was done with chlorine. Bromine does not bleach as well as chlorine, although there is some effect.

5.4 To Prepare and Study Iodine

The precautions with respect to fumes apply here as well, although not to such a great degree. If the watch glass has ice kept in it and if the burner flame is kept low there should be no iodine escaping from the beaker. The best results are obtained when the flame is kept very low, thereby allowing the iodine produced to sublime onto the cool watch glass. Be sure to leave the watch glass in position for several minutes after the heating has stopped. Observe the precautions for handling the carbon tetrachloride.

Probably the most important part of this sequence is found in the two questions at the bottom of page 281. Be sure that when you talk about position in the table you relate back to the active metals studied in the first two experiments.

5.5 Relative Activity of the Halogens

As with the previous experiments, caution your students about the dangers of the vapours involved in this experiment – especially that of the carbon tetrachloride. And caution them not to use their thumbs as corks!

If you wish you can go back into the previous labs and get the answers to parts of questions 1, 2 and 4 of experiment 5.1 as part of your pre-lab.

We have found it useful to compare the relative activities by considering these reactions as a competition for electrons. In the case of the halogens the ions are the "haves" and the atoms are the "have nots". When an atom and an ion are thrown in the test tube together the particle that ends up as the ion is the more "powerful" electron grabber. For example, the chlorine atom is able to take electrons from the bromine ion. This indicates that it is a better electron grabber than bromine. If you have not already done so, this is a good place to make the point that no atom or ion *gives* electrons away. They are always taken away by some other substance. In a redox reaction there is always a competition between two substances.

You should also make the point that although bromine and iodine are not in the same class as chlorine as far as the ability to take electrons is concerned, they are still far ahead of the metals. Point out that this ability to reduce is a relative thing.

5.6 Reaction of Sulphuric Acid with Sodium Chloride, Sodium Bromide and Sodium Iodide

As before it is very important that you treat the three parts of this experiment together. At the end of the three parts the students should see the similarities in the behaviours of the three halogens. Remember what we are aiming for: we want the student to see that the periodic table is arranged in families of elements. Each family is recognized by similarities in chemical and physical properties. At the same time your students will be introduced to the properties of the hydrogen halides.

Take the usual precautions with respect to fumes.

Once you get an answer to the last question on page 286 you can show some of the properties of hydrogen fluoride by doing the following demonstration.

Cover a glass plate with a thin layer of wax. Use a small screwdriver to remove some of the wax in the form of a design – an initial, for example. In a fume hood mix sodium fluoride and concentrated sulphuric acid. Place the glass plate over the beaker with the waxed side down and the design directly over the reacting mixture. Leave the apparatus this way overnight. Next day, remove the wax from the plate by heating in hot water (this is messy), and observe your initial etched in the glass! This is how artists etched glass for many years. It also shows why one does not keep hydrofluoric acid in a glass bottle but in a wax bottle. One last word: when you wash out the beaker you will probably discover how one could go about "frosting" the inside of a container.

5.7 Test for the Chloride, Bromide and Iodide Ions

The principal objective here is for the student to learn a method for identifying the three ions: chloride, bromide and iodide. We do this, as indeed we identify anything, by using a unique set of properties.

Part A

If you would like to make this a more extensive test then have your students add silver nitrate to samples of bromide and iodide. The precipitates that form in all three cases are slightly different in colour. Have your students expose all three precipitates to the sun or at least a bright light and they will see the beginnings of the chemistry of photography. If you wish to carry this further go straight to experiment 7.6 on page 365.

The completion of your discussion of this lab is a convenient time to give your students a lab test. Have available all of the reagents from experiments 5.2 to 5.7. Prepare at least three unknowns for the class to try to identify. We suggest that one of them be sodium bromide in solution, one be sodium sulphate in solution and one be a mixture of sodium chloride and sodium iodide.

5.8 The Properties of Third-Row Elements

This lab provides a good opportunity for students to suggest their own lab exercises. Listen to all of their suggestions and allow any student with reasonable ideas to try them out. You should also have a copy of a handbook so that those who want to can look up other data. Try to encourage your students to consider as many different things as possible. You may even want them to graph some of the numerical data as a preliminary exercise to the graphical analysis that follows.

You may wish to spend a few moments talking about those substances we call amphoteric hydroxides. These are the hydroxides such as $Al(OH)_3$ that exhibit the properties of an acid when they are allowed to react with a strong base. They represent an interesting transition between the acidic oxides and the basic oxides as we move across the periodic table.

5.9 Silicon

This is a really good opportunity for your students to practice the discovery method of science. Of course much depends on the materials you set out for them. If there happens to be a conductivity tester available you will be surprised at how many of your students will think of trying it!

You should have available dilute acids and dilute bases, indicators, balances, limewater and any other substances that your students have made use of over the course of the term.

Graphical Analysis

When your students describe the general shape of these graphs, try to get them to tell you that they go "up and down, up and down, up and down." Work towards the idea that they look like a wave form, and that will lead you to an introduction of the term that describes waves or any other repeated motion – *period* – and hence *periodic*.

The meaning that we want to attach to the term periodic is "repeating at regular intervals." Since each of these graphs represents a plot of some property, then we should be able to say that the properties of the elements repeat themselves at regular intervals. They are periodic. Since the table we have been working on is based on these repeating properties we have a "periodic table of the elements."

The last exercise is rather like assigning electronegativity values à la Pauling. What we are really trying to do is consolidate the student's ideas on metallic and non-metallic properties. The student may need a lot of help here but we think that it is worth the effort. It certainly can be confusing if one is not careful!

What follows is a question you may wish to use in your evaluation of this unit.

Unit 5: Sample Test

After the Space Wars had ended Luke Skywalker returned to his home galaxy to resume his studies. The same laws and theories apply on Tatooine as apply on earth. Unfortunately all the elements had different names than the ones Luke learned on earth. Using the information provided, try to match the Tatooine name for an element with the earth name of the same element.

1. Alpha is a black non-metal that does not conduct electricity.
2. Epsilon is an inert gas. One litre of this element has a mass close to 1 g.
3. Lambda reacts vigorously with water and burns with a yellow flame.
4. Gamma is very similar to lambda in appearance and properties but has a larger atomic mass.
5. Beta is a very reactive gas with six electrons in the valence shell.
6. Kappa has similar chemical properties to beta but is a solid.
7. Tau has two valence electrons, reacts slightly with hot water and burns vigorously in air.
8. Pi is chemically similar to epsilon but is only about one fifth as dense.
9. Theta is similar to alpha but appears to be more metallic in its appearance.
10. Omega is below tau in the periodic table.

Unit 6
Organic Chemistry

Because carbon is unique among the elements and because it is so versatile we have included a unit on the study of carbon and some of the compounds derived from carbon. The unit simply grazes the surface of this enormous branch of chemistry, but it does introduce students to the topic and to the vastness of the subject. In this unit the students will first find that the element appears in more than one form in nature. Further, they will find that carbon has the ability to bond covalently with other carbon atoms to form endless chains or ring-type molecules. This unique property of carbon means that there are many more molecules of matter that contain carbon than there are molecules that do not contain carbon. Students find that the enormous amount of material to be learned in organic chemistry is made somewhat simpler by dividing molecules into two parts, namely the unreactive hydrocarbon part and the much more reactive functional group.

Molecules with the same functional group are placed in the same family of compounds. Each member has some different properties than other members but many properties are common to all members of the family.

The students also discover that high molecular mass carbon compounds are solids and that low mass molecules can be joined to form these most useful solids called plastics.

In this unit students will become familiar with the systematic naming of organic compounds for alkanes, alkenes, alkynes, alcohols, aldehydes, ketones, organic acids and esters. By the end of the unit students should also be familiar with certain reactions of these organic compounds and with the functional groups of these compounds.

6.1 Heating Substances out of Contact with Air

The most difficult part of this experiment is setting up the apparatus illustrated on page 311. A simpler procedure is to heat the materials in a covered crucible, as in experiment 6.3. This is less time consuming and is readily available for all students to use.

The intent of the exercise is to illustrate that things that were once living have a basic framework that is similar in each different substance used. Most students will recognize the framework as being constructed of carbon atoms, hence an introduction to the unit.

We find it useful to heat a crystal of bluestone under the same conditions as the other materials. This serves as a control because it does not originate from living matter, and when heated it will not become black and it still retains its shape. The bluestone aids in establishing the idea of organic chemistry or the idea of carbon chemistry.

6.2 Carbon in its Different Forms

This exercise is simply a continuation of experiment 6.1. When the materials are heated out of contact with air, students believe the black residue is carbon from previous experiments with the element. Experiment 6.2 simply illustrates that when the black residue is crushed into a powder from the various substances, it is virtually impossible to recognize the source of the black powder. The point to be established is that the basic structure of the various materials heated is formed from the same material – carbon.

The Carbon Atom

Now that the unit has been introduced and students are aware that the unit concerns itself with the study of carbon and carbon compounds, we find this a convenient place to discuss the versatility of the carbon atom. At this point describe to the students the unique bonding features of carbon and indicate why its study has become such a vast and significant branch of chemistry. The *significance* of carbon and its compounds will introduce experiment 6.4.

6.4 Carbon Compounds from Crude Oil

We do this experiment mainly because it is so

relevant, however in the process the ideas behind fractional distillation may also be introduced. In addition, since petroleum is a complex mixture of *numerous* liquid hydrocarbons it provides a reason for studying and talking about hydrocarbons and their formulas and names.

We have always been able to obtain petroleum (free of charge) from the major oil companies. Here are two addresses we have used, but there may be others closer to your area.

Texaco Canada Ltd.
1425 Mountain St.
Montreal, Quebec

B.P. Oil Refinery
Oakville, Ontario

Because there will be some hydrocarbon vapour in the air there may be some danger, so we advise that every safety precaution be taken. We have done this experiment on many occasions without incident. Students should of course wear safety glasses and an asbestos shield should be placed between the burner and the collecting vials. Iced water in the beaker into which the various fractions are being collected will help considerably in the condensation of the vapours produced by the heating of the petroleum.

The post-lab discussion to this experiment provides a convenient time to introduce the naming system for organic compounds, in particular the alkanes.

6.5 To Prepare and Study the Properties of Methane

We do this experiment in order to have students work with an alkane and also to illustrate that all hydrocarbons are *not* liquids, as indicated by experiment 3.4. This experiment is potentially *dangerous* and as is usual with every experiment of a dangerous nature, every safety precaution must be exercised. When procedure 4 is reached it is absolutely essential that the gas produced ignite *quietly* in the test tube before the jet is ignited. At this point a fairly high concentration of methane is being produced, so the jet can be ignited with a reasonable degree of safety.

Since the gas is very much like hydrogen in observable properties to the end of procedure 4 in the experiment, we have the students collect some of the product of combusion (procedure 5) and then react it with limewater to indicate the presence of carbon dioxide.

6.6 Alkanes and Alkenes Compared

The reactions in this experiment help to define certain classes of organic compounds, namely unsaturated and aromatic.

It may help to have available ball-and-stick models of the hydrocarbons used in this experiment. These may be constructed ahead of time and left on a side counter out of the way of the lab area, so that students can examine them as they proceed with the experiment.

The hydrocarbons used here are flammable, which students should be made aware of ahead of time. The fumes from these compounds are quite distinct and may be unpleasant to some people. Lab fans help to alleviate this problem.

Dispensing organic compounds from small bottles will eliminate the problem of having large quantities of the compounds open in the lab.

If you find that some of the results of this experiment don't agree with your thinking or with what you expected, don't be alarmed. Apparently it is virtually impossible to avoid contamination of some of these hydrocarbons. If the bromine, for example, decolourizes when you did not expect it to, it will probably be due to such contamination. These reactions with contaminants will occur gradually over a period of five to ten minutes.

6.7 Preparation and Properties of Acetylene

This experiment is intended to introduce to the students the idea of triple bonds as well as the alkyne family of organic compounds.

Because acetylene is so combustible we generally do this experiment as a demonstration.

The various amounts of gas collected and burned illustrate the various degrees of combustion that may result, depending upon the ratio between volume of gas and volume of air present.

The incomplete combustion of acetylene produces carbon dioxide and carbon.

$$2C_2H_2 + 2O_2 \longrightarrow CO_2 + 2H_2O + 3C$$

As the oxygen supply is increased, acetylene burns with a more brilliant and a much hotter flame, producing only carbon dioxide and water.

$$2C_2H_2 + 5O_2 \longrightarrow 4CO_2 + 2H_2O$$

As the calcium carbide reacts with the water, a solution of calcium hydroxide is formed. If some of the solution is filtered, passage of carbon dioxide through it will turn it a milky colour.

Names of Large Molecules

Before leaving the study of hydrocarbons we find it

convenient to establish the naming system for branched molecules. This need not be done here but could well be done after the students become familiar with several functional groups and the role of these functional groups in the naming system.

6.8 Making Ethanol by Fermentation

This experiment is certainly not essential to the unit, but it does provide a convenient way to introduce alcohols and functional groups. Further, this is a chance to illustrate the use of a condensor or of ground-glass equipment and heating mantles if they are available.

We have not had great success with burning some of the distillate, which has made the experiment somewhat disappointing.

6.9 Some Reactions of Alcohols

This experiment is intended to illustrate the oxidation of an alcohol under various conditions. In an acid solution or in a basic solution the reactions are quite evident, as indicated by changes in the colour of the permanganate.

Acid

$$5C_2H_5OH + 4MnO_4^- + 12H^+ \longrightarrow 4Mn^{+2} + 11H_2O + 5CH_3COOH$$

Base

$$C_2H_5OH + 4MnO_4^- + 4OH^- \longrightarrow 4MnO_4^{-2} + CH_3COOH + 3H_2O$$

The reaction in a neutral solution may be so slow that it goes unnoticed by the students.

The reaction with the alcohol and copper is certainly interesting and worthy of demonstration at least. When the wire is heated it turns dark (from the oxide formation and perhaps a bit of carbon). When the hot wire is placed over the alcohol it turns alternately shiny and dark. A very distinct odour different from the alcohol will be noted. Students often recognize it as formaldehyde.

$$CuO + CH_3OH \longrightarrow \underset{\text{Formaldehyde}}{HCHO} + H_2O + Cu$$

6.11 Mixing Alcohols and Acids

The experiment gives very good results from a simple reflux apparatus. The only problem one ever encounters with this experiment is that drops of liquid form in the reflux tube and are forced out of the top of the tube as the pressure increases in the flask. This problem can be overcome fairly well by using a glass tube with an inner diameter of about 4 mm. The larger tube allows the vapours to condense and run back into the flask without collecting and forming large drops or beads that will block the tube.

These reactions are catalyzed by mineral acids or bases. However, acids like HCl or H_2SO_4 are generally used.

6.12 Preparation of an Ester

After students have prepared an ester and have some understanding of the chemistry involved, we collect a number of alcohols and acids and assign each lab group a different combination of alcohol and acid. The esters prepared by each lab group may be collected and labelled for examination by the other students in the class.

Ethyl acetate produced from ethanol and acetic acid is characterized by a distinctive fruity odour that is familiar to most of the students.

Ethyl butyrate produced from ethanol and butyric acid has a distinctive pineapple odour. This ester is especially interesting because of the unpleasant odour of butyric acid.

Methyl salicylate is produced from methanol and salicylic acid catalyzed by a very small amount of sulphuric acid. This ester has an odour resembling that of oil of wintergreen.

Amyl acetate from acetic acid and amyl alcohol has the pleasant odour of banana oil.

6.13 Condensation Polymers

A unit of organic chemistry would not seem complete without some brief mention of polymers and polymerization. However, experiments suggested all involve the production of condensation polymers, because we have been unable to find an addition polymer that can be produced with reasonable ease and safety in a high school laboratory.

6.13 Part A

Because of the clean-up difficulty we have found that the use of a disposable container is best for this experiment. With gentle heating a small baby food jar works well. Hospitals use a disposable baby formula bottle which we have used for this experiment as well.

The success of the exercise depends essentially upon the heating, which should be done gently until large bubbles puff up through the mixture, much like bubbles break through a thick porridge as it is cooking.

The odour and smoke produced from this experiment may cause discomfort for some people, so exhaust fans should be used.

6.13 Part B

This experiment produces a beautiful pure white polymer when the experiment is successful. The secret of success is to be certain that the urea-formaldehyde solution is *saturated.*

6.13 Part C

The production of nylon in this experiment has been called the rope-trick experiment, because from a *small* beaker, metres of nylon may be drawn in a continuous ribbon. This experiment is simple to perform as long as the water solution is poured in so as not to mix with the bottom layer. The demonstrator of this experiment should wear rubber gloves.

Unit 6: Sample Test 1

1. Name these substances:

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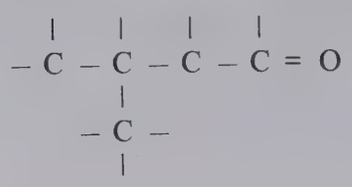

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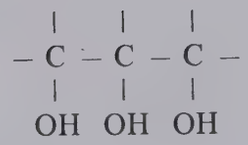

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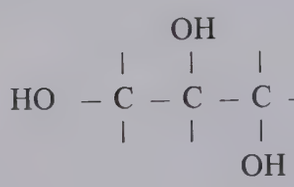

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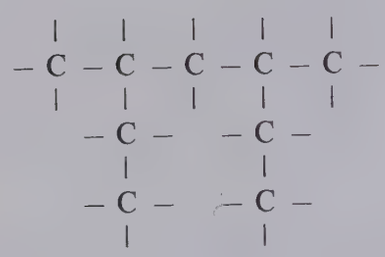

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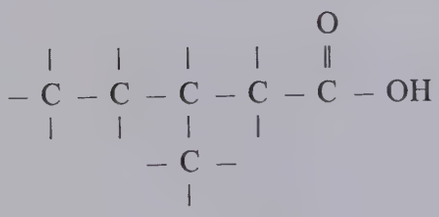

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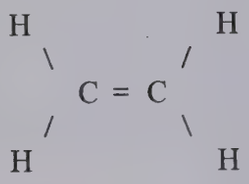

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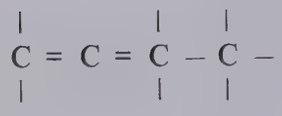

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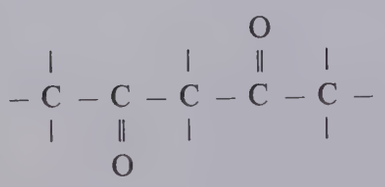

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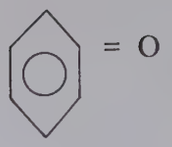

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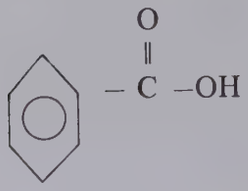

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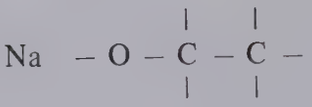

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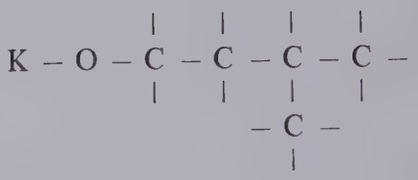

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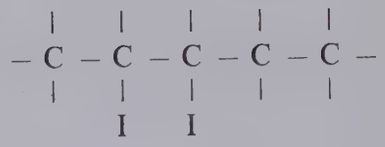

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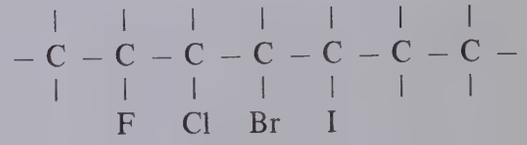

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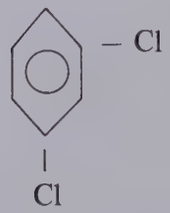

Unit 6: Sample Test 2

1. Draw the structural formula for:
   (a) 2-2-4-trimethyl pentane
   (b) any iso heptane
   (c) 3-octanone
   (d) 1-2-diiodocyclobutane
   (e) 3-chloro-4-bromo-5-iodo-1-cyclopentene
   (f) propanedial
   (g) 1-6-hexandiol
   (h) 1-6-hexadiol
   (i) 1-3-benzene dioicacid
   (j) tert-hexanol
   (k) Draw *all* the isomers of hexane and name them.
   (l) Show how ethandioic acid would polymerize with 1-3-propandiol. Name the type of polymer in *two* ways.
2. For each reaction select the structure of the organic product from the second sheet and place its number in the bracket. Name the product in the space provided on this sheet. Hand in both sheets.
   (a) 2-pentene + I_2 $\longrightarrow$ _____ ( )
   (b) n-hexanol + K $\longrightarrow$ _____ ( )
   (c) n-hexanol + HCl $\longrightarrow$ _____ ( )
   (d) formic acid + Na $\longrightarrow$ _____ ( )
   (e) ammonium butanate $\longrightarrow$ _____ ( )
   (f) butane + [O] $\longrightarrow$ _____ ( )
   (g) ethanol + acetic acid $\longrightarrow$ _____ ( )
   (h) propanoic acid + LiOH $\longrightarrow$ _____ ( )
   (i) tert-butanol + [O] $\longrightarrow$ _____ ( )
   (j) methamine + methanoic acid $\longrightarrow$ _____ ( )

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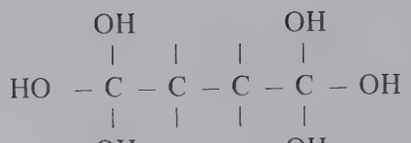
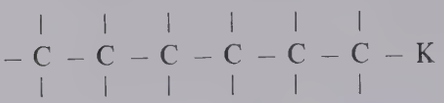
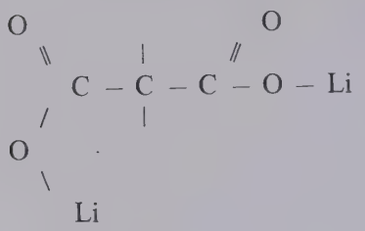
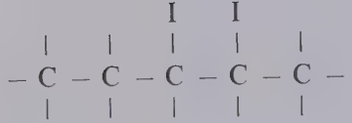
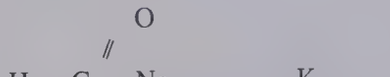
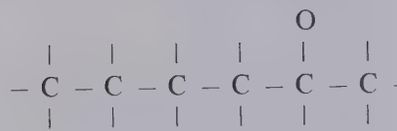
1. 

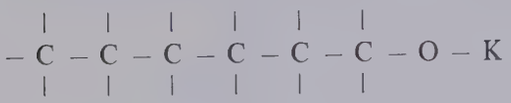
2. 

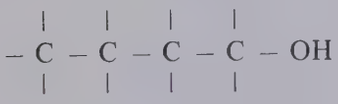
3. 

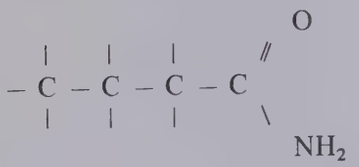
4. 

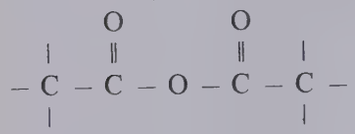
5. 

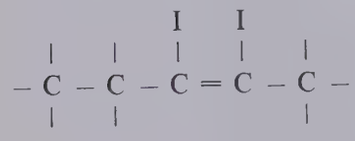
6. 

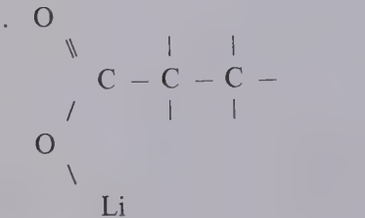
7. 

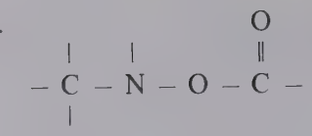
8. 

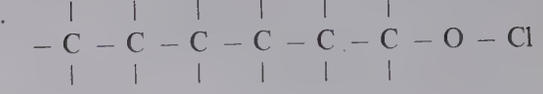
9. 

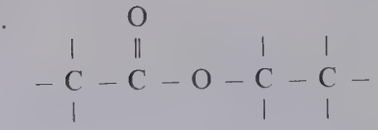
10. 

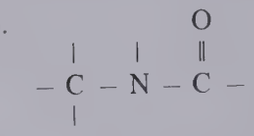
11. 

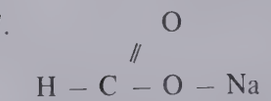
12. 

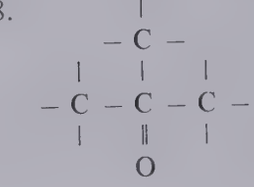
13. 

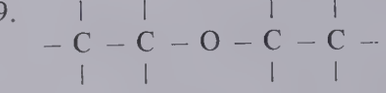
14. 

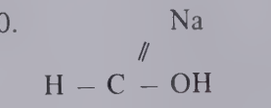
15. 

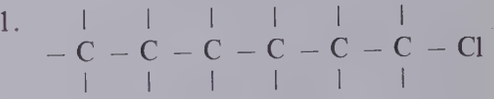
16. 

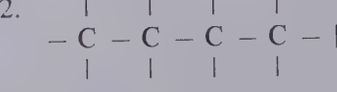
17. 

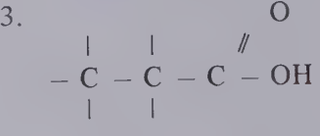
18. 

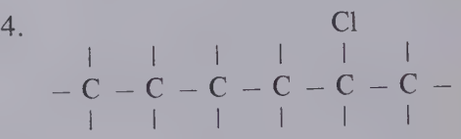
19. 

20. 

21. 

22. 

23. 

24. 

25. inorganic

Unit 7
The Marketplace — Applications of Chemistry

This unit serves as an introduction for teachers and students to some of the topics that fit under the heading of consumerism. It is our intention here to introduce a number of relevant topics that may be pursued further with additional experimentation or research. For this reason the unit may seem a little disjointed, with a series of unrelated topics.

Because a knowledge of equations, formulas of binary compounds and acids and bases is required for Unit 7, we suggest that best results are obtained only after Unit 4 has been completed.

The first five experiments in Unit 7 all relate to water hardness and topics such as soaps, detergents, bleaches and how odours or fragrances are given to soaps.

Other topics that are examined in Unit 7 are photography, metals and their properties and qualitative analysis. You may wish to investigate only one of these topics, or all of them.

7.1 Properties of Hard Water

This experiment is intended to demonstrate the difference between permanent and temporary hardness in water as well as ways of solving the hardness problem in water.

Hard water contains dissolved matter such as calcium bicarbonate or calcium chloride. Hardness is generally classified as permanent or temporary depending on whether it may be removed by boiling or not. Procedure 5 demonstrates the removal of hardness by boiling. In this case the soluble bicarbonate will decompose, yielding the carbonate that will precipitate removing the calcium ion from solution.

$$Ca(HCO_3)_2 \longrightarrow CaCO_3 + CO_2 + H_2O$$

Procedure 8 of this experiment illustrates the removal of permanent hardness. Generally hardness of this sort is removed by the addition of some chemical that will precipitate the ion(s) causing the hardness in the water. For example, if the hardness is caused by calcium ions then the addition of sodium carbonate (which is soluble) to the water would cause the calcium to precipitate as calcium carbonate.

$$CaSO_4 + Na_2CO_3 \longrightarrow \underset{\text{(precipitate)}}{CaCO_3} + \underset{\text{(soluble)}}{Na_2SO_4}$$

7.2 Preparation of Soap
7.3 Properties of Soap

The oil used for the preparation of soap does not necessarily have to be coconut oil but may be any oil that you have available in the lab.

lauric (coconut oil) R = 12
linoleic (cottonseed oil) $C_{18}H_{32}O_2$
linolenic (linseed oil) $C_{18}H_{30}O_2$
oleic (peanut oil) $C_{18}H_{34}O_2$

A 5M NaOH solution of sodium hydroxide is potentially dangerous and every safety precaution should be followed, since spattering of the oil-base mixture may occur.

The heating should be continued until a thick, gelatinous mixture results, but with some oils this may not occur until the mixture is cooled. If after a 15-min heating period the mixture is not thick and sticky, simply allow the mixture to cool.

Since the NaOH will be in excess it will be absolutely essential that the washings of the prepared soap be thorough to remove the NaOH. The last of these washings should be done with a salt solution.

Experiment 7.3 illustrates the ability of the soap to:

1. produce a suds with soft and hard water (procedure 2);
2. affect the surface tension of water (procedure 3);
3. suspend oil droplets throughout water (procedure 6).

7.4 Bleaching Action of Chlorine

The demonstration at the start of experiment 7.4 illustrates the production of oxygen by the reaction between chlorine and water.

$$Cl_2 + H_2O \longrightarrow HCl + HClO$$

Hypochlorous acid, being rather unstable, decomposes to yield atomic oxygen, which ultimately is the

cause of the bleaching action of the coloured material.

The acid from Part B simply increases the production of hypochlorous acid. As the concentration of HClO is increased so is the concentration of atomic oxygen. Since the oxygen causes the bleaching, a high concentration of it will make the bleaching more evident.

Chlorine and water produce hypochlorous acid and hydrochloric acid.

$Cl_2 + H_2O \longrightarrow HClO + HCl$

But HClO is unstable:

$HClO \longrightarrow HCl + [O]$

$[O] + \text{dye} \longrightarrow$ colourless compounds

7.5 How Fragrances are Made

See comments on experiment 6.11, "Mixing Alcohols and Acids."

7.6 Photographic Film

This experiment is certainly of interest to the students and provides a slight change from the normal routine of performing experiments.

In the preparation of the film, a layer of light-sensitive silver bromide is formed on the filter paper. Best results are obtained with a thin, uniform layer of the silver bromide.

$NaBr + AgNO_3 \longrightarrow AgBr + NaNO_3$

The cardboard sandwich serves as a camera. This should be prepared *before* the negative, otherwise it will have to be prepared in a darkened room.

To expose the film we have found that best results are obtained by allowing the students to go outside if the sun is shining. Failing this, several bright lights should be available in the classroom. The regular room lights are hardly adequate.

7.7 Gardener's Sclerometer

This experiment is a simple means of testing the hardness of various metals. The idea was developed by a colleague named Gardener, thus the title of the experiment. The height of the bounce of the nail is proportional to the hardness of the metal. Further, the harder the metal the more difficult it will be to mark or scratch the surface.

7.9 Flame Tests

A solution may not be adequate to provide enough salt to produce a significant change in the colour of the flame. If this is the case, form a small loop in the end of the wire and dip it directly into a sample of the dry salt. However, the sample obtained should be small to prevent particles of the salt from falling into the burner, which may interfere with future flame tests.

The colours produced by the common metals are:

barium – pale green
strontium – crimson
calcium – orange
sodium – yellow
potassium – violet
lithium – pink

7.12 Qualitative Analysis 3

To test for nitrates by the brown-ring test is certainly of interest and is a requirement for experiment 7.13. However, the handling of concentrated sulphuric acid by a class of students is potentially dangerous. We have prepared acid-dispensing boxes, which remain in fixed locations in the laboratory, and from which acid may only be dispensed by a dropper.

Success in this experiment depends upon a *freshly* prepared solution of ferrous sulphate and careful adding of the acid so that the acid and the remaining solution do not dissolve into each other.

In the phosphate test the heating should be done gently in a hot water bath. As the solution approaches about 50°C, a yellow precipitate begins to form in a phosphate solution.

7.13 Testing Fertilizers for Elements Necessary for Plant Growth

Experiment 7.13 may be used as a test of the students' ability to apply what they have already learned about qualitative analysis, since flame tests as well as the brown-ring and phosphate tests are applied. Of course these tests have already been developed in previous experiments. It may be of interest to provide several samples of commercial fertilizers, but in your selection choose one or two samples that are lacking in one of the three normal constituents of fertilizers. Our school is in a rural setting and this experiment has evolved into a mini unit in itself, with students providing various samples that they bring from home.

Lab Test

The conclusion of a section on qualitative analysis provides an excellent opportunity for a lab test. Provide the students with an unknown made up to your specifications. The job of the students will be to carry out the necessary tests and to identify the constituents in the unknown. Students generally find this type of exercise interesting and challenging.

Unit 8
Laboratory Skills

It seems to us that besides teaching something about the content of science and about the processes of science, any good science course must also concern itself with attitudes, values and some very basic skills.

It is almost impossible to teach the first two from a textbook because they depend so much on the attitudes and the values of the teacher. It is also difficult to teach skills from a text.

It is our feeling that most teachers will want to teach lab skills in their own way and pass on the specific techniques they think are best. For that reason we chose not to include these skills within the main body of the text. Instead we have included them as a supplement, which you may refer to under several different sets of conditions.

1. These may be used as pre-labs, before you teach the particular skill in your own way.
2. You may wish to use them to help students who were absent when you covered a particular skill.
3. When your class is doing a lab you may see a student who is lacking in a particular skill. The text may be used as the first step in helping the student.
4. There is always the possibility that you may want your class to learn the skills directly from the text.

Each of the skills included in this unit is needed at some point in the course. They appear in roughly the order that they are needed, assuming that you follow our unit order.

We do the exercise on the bunsen burner as a lab. We do the same with the balance exercise.

The points on glassware are simple and straightforward. Most students are not aware of steps 3 and 4.

Filtration is an important skill. Every student should be able to fold a filter paper properly.

The volume-measurement section should be familiar to most students, except for the pipette and the burette.

We recommend that you spend a day on the glass manipulation exercise just because students enjoy it so much. Be sure you use soda glass and that you have an asbestos mat for each student to put the hot glass on.

We can't underestimate the importance of the section on handling liquids.

For many teachers the section on identifying odours will be new material. We were only made aware of this technique a few years ago and it has been our experience that most teachers have never learned the proper method for checking on odours.

The melting point determination is a very useful technique all students should be aware of. The same can be said for the section on heating in a crucible.

Evaluation for this unit should be done entirely in the lab. Once a technique has been mastered, students must be aware that they are expected to do it properly whenever the need arises. Failure to do so should result in a penalty of some sort.

A common method of handling this section is to award each student 25 marks for technique at the start of the year. Every time you see bad or sloppy technique deduct one or more marks from that total. This only works if you tell the student about each deduction and the reason for it.

Unit 9
Independent Study Projects

This unit may make many teachers apprehensive. It need not do so. It does represent a lot of work for you but it can be very rewarding. If you believe, as we do, that your students should be given a chance to do some real science then this unit is ideal for you.

Before we start, let's have an understanding. All year long you have taken the time to help out the slower students. You have held up lessons and you have explained some points six times to seven different people. In short, you have tried very hard to give every consideration to those who are having trouble. Now it's time to give those students who have not had any difficulty their reward. If you do, it must be with the certain knowledge that some students will not get a great deal out of doing their projects. Do this for the other ones!

Before you begin:

1. Make sure your library has a good supply of reference material. The books listed on page 436 have been very useful to us.
2. Start your preparations at least two months in advance of the day when students will actually begin their projects. On this page and the next you will find a timeline that can serve as an example of the preparations you should make.
3. Prepare your students for the basic concept of the project well in advance. They should understand that this is going to be their chance to practice some real science. They will chart their own course as they try to solve some problem in a scientific way. In order to begin properly, each student should be looking for a topic that is of interest to him or her. Be sure they understand that once they have identified an area of interest they must come up with a question or problem that needs an answer.

   Be sure to go over the basic plan of attack one more time.

   (a) Gather.
   (b) Organize.
   (c) Evaluate.
   (d) Interpret.
4. Outline well in advance the kind of report that will be acceptable to you. We suggest that you be as flexible as possible here. Remember that the basic idea is for students to solve problems on their own. Each project is not just another vehicle for a formal lab report. Indicate that you will accept tapes, films, photographs, demonstrations – any method that will permit students to communicate their ideas to you. All they should really have to prove is that they learned something from the experience.

Project Timeline

Now for the timeline. We suggest that you plan on setting aside five days of class time. During this period the students will work independently of each other, each on his own project. You will be in the lab ready to offer advice and talk over problems. We have found that late May is the best time for these projects.

Once you have picked the dates for the five days assigned to the projects back up 40 teaching days on the calendar to what we will call day 1 of the timeline.

Day 1

Introduce the project. Outline how things are going to work. Establish the timeline for all students. Make certain they understand that all dates must be met unless they have made special arrangements in advance. Give each student a copy of the timeline. They should read page 402 in the text and then scan the projects listed in Unit 9 from pages 403-435. They should also read pages 435-436.

Day 6

All students hand in a piece of paper on which they have written the title of their projects and a statement of the problem or question they will try to answer. If any of these are not acceptable to you, be sure to contact the student before the next hand-in is due.

Day 11
Students hand in reports in which they outline very briefly how they intend to find the answer to their questions. Detailed experimental outlines are not required here.

Day 16
Students hand in a sheet of paper on which they have worked out the procedure for any experiments that they intend to do. These should be in enough detail for you to determine how safe the experiments will be and what materials are needed. If you do not have some of the materials listed, confer with the student to determine how he intends to provide these things.

It is probably worth noting here that while we tell the students that we will not provide anything except the normal lab materials, we do go out of our way to help them find the things they need for their projects. Sometimes you can offer suggestions for substitutions or minor changes in procedure to make use of the materials that you have available.

Try to find time over the next two weeks to talk to every student about his or her project. Make an effort to encourage all of them to be original in their thinking and their methods of presentation.

Day 26
Students hand in a detailed list of the materials and apparatus they intend to use in their projects. This is almost your final chance to veto anything you don't like or to suggest changes in procedures. Look over the list carefully from the standpoint of lab safety.

Day 31
Students hand in a brief statement explaining how they plan to report the results of their projects to you. Their reports should include a list of all the resource materials they have found and which of these they intend to draw upon.

Day 36
Students hand in timelines showing how they plan to budget time over the five days assigned to the projects. This report should not only tell you when the student plans to do what but also where; and in some cases with whom! You will find that most of them will greatly overestimate how much they can accomplish in the five-day period.

During this last interval you should talk to the ones who have obviously taken on too much. Persuade them to scale their projects down to a more reasonable size.

Day 40
This is it! Prepare to get swept off your feet! You will find this a very hectic period of the year. Try not to get too angry at the students who seem to be wasting time; remember that they have to learn to work on their own sometime. Just make a mental note so you can refer to their behaviour at a later date.

Since each project requires a lot of preparation, we feel that it should be worth a fair number of marks. We suggest that you mark the project out of about four times what you would give for a normal class test. Don't forget that each step in the preparation should be worth something!

Try not to be prejudiced in favour of the formal written report. The basic idea is that students are trying to communicate their findings to you.

Good luck!

Answers

Unit 1: Sample Test 1

1. (a)
2. (b)
3. (c)
4. (b)
5. (c)
6. (b)
7. (c)
8. (c)
9. (d)
10. (d)
11. (c)
12. (d)
13. (a)
14. (c)
15. (c)

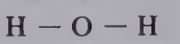
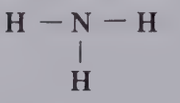
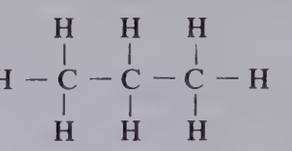
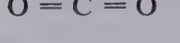
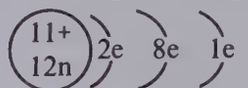
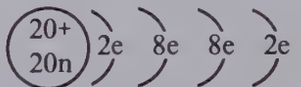
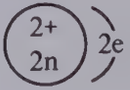
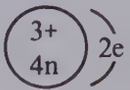
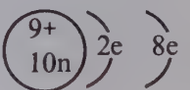
Unit 1: Sample Test 2

1. (b)
2. (c)
3. (e)
4. (d)
5. (a)
6. (a)
7. (b)
8. (c)
9. (b)
10. (a)
11. (c)
12. (c)
13. (a)
14. (d)
15. (b)
16. (d)
17. (d)
18. (c)
19. (e)
20. (e)
21. (a)
22. (d)
23. (c)
24. (e)
25. (a)

Unit 2: Sample Test

1. (a) CS_2
   (b) SO_2
   (c) P_2O_5
   (d) SF_6
2. 44.0 g, 64.1 g
3. 28 g
4. 0.03 mol
5. (a) potassium
   (b) calcium
   (c) cobalt
   (d) carbon
   (e) lead
6. (a) hydrogen iodide
   (b) boron trichloride
   (c) carbon tetrachloride
   (d) oxygen dichloride
7. (a) 0.67 mol
   (b) 0.03 mol
   (c) 2.0 mol
8. 5.3×10^{-6} g
9. (a) 4.2×10^{24} atoms
   (b) 3.0×10^{24} atoms
   (c) 5.4×10^{24} atoms
   (d) 1.26×10^{25} atoms
10. 6.02×10^{23} molecules
11. 10 mol
12. (a) 0.01: 0.01: 0.01
    (b) 1: 1: 1

Unit 3: Sample Test

1. (a)
   (b)
   (c)
   (d)
   (e)
2. (a) 
   (b) 
   (c)
   (d) 
   (e) 
3. (a) $Mg + S \longrightarrow MgS$
   (b) $C + O_2 \longrightarrow CO_2$
   (c) $NaBr \longrightarrow Na^+ + Br^-$
   (d) $2PbO_2 \longrightarrow 2PbO + O_2$
   (e) $Na_2S + CuF_2 \longrightarrow CuS\downarrow + 2Na^+ + 2F^-$
4. $3Na_2CO_3 + 2H_3PO_4 \longrightarrow 2Na_3PO_4 + 3H_2O + 3CO_2$
5. (a) 
   (b) 
   (c) 
   (d) 
   (e) 

6.

Chlorine	*Sulphur*	*Nitrogen*	*Carbon*	
$NaCl$ sodium chloride	Na_2S sodium sulphide	Na_3N sodium nitride	Na_4C sodium carbide	sodium
$MgCl_2$ magnesium chloride	MgS magnesium sulphide	Mg_3N_2 magnesium nitride	Mg_2C magnesium carbide	magnesium
$AlCl_3$ aluminum chloride	Al_2S_3 aluminum sulphide	AlN aluminum nitride	Al_4C_3 aluminum carbide	aluminum
HCl hydrogen chloride	H_2S hydrogen sulphide	H_3N hydrogen nitride (ammonia)	H_4C hydrogen carbide (methane)	hydrogen

Unit 4: Sample Test 1

1. (c)
2. (c)
3. (b)
4. (b)
5. $HMnO$
6. H_2TeO_5
7. (a) $HNO_3 + KOH \longrightarrow KNO_3 + H_2O$
   (b) $2H_3PO_3 + 3Ba(OH)_2 \longrightarrow 6H_2O + Ba_3(PO_3)_2$
   (c) $HClO + NH_4OH \longrightarrow NH_4ClO + H_2O$
8. $H^+ + OH^- \longrightarrow H_2O$

Unit 4: Sample Test 2

1. $2KClO_3 \xrightarrow[\text{catalyst}]{MnO_2} 2KCl + 3O_2$
2. H_2CO_3 carbonic acid
3. any other acid properties
4. A titration is a procedure used to determine the concentration of an unknown substance. The pipette and burette are used to measure liquid volumes. The indicator is used to determine the end point.
5. (a) $H_2 + CuO \longrightarrow H_2O + Cu^0$
   (b) $2H^0 - 2e \longrightarrow 2H^+$
   (c) $Cu^{+2} + 2e \longrightarrow Cu^0$
6. $2HNO_3 + Li_2CO_3 \longrightarrow CO_2 + H_2O + 2LiNO_3$
7. (a) hydrogen bromide
   (b) nitrous acid
   (c) calcium sulphate
   (d) per phosphoric acid
   (d) sodium sulphite
   (f) barium hydroxide
   (g) iron III chloride
   (h) hypochlorous acid
   (i) aluminum per bromate
   (j) cesium hydroxide
8. (a) Na_3PO_4
   (b) NH_4OH
   (c) $HClO$
   (d) KCH_3COO
   (e) $LiClO_4$
   (f) H_2S
   (g) HIO_3
   (h) $NaHCO_3$
   (i) $Sr(NO_3)_2$
   (j) $BaSO_2$

Unit 5: Sample Test

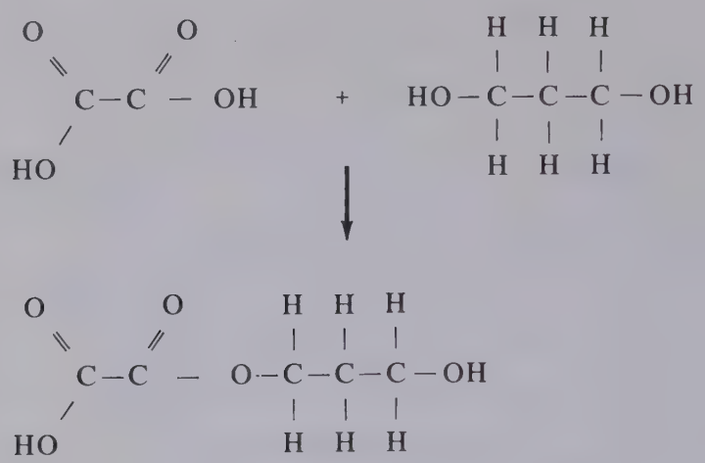
1. carbon
2. neon
3. sodium
4. potassium
5. oxygen
6. sulphur
7. magnesium
8. helium
9. silicon
10. calcium

Unit 6: Sample Test 1

1. (a) 3-methylbutanal
   (b) 1-2-3-propantriol
   (c) 1-2-3-propantriol
   (d) 3-5-dimethylheptane
   (e) 3-methylpentanoic acid
   (f) ethene
   (g) 1-2-butadiene
   (h) 2-4-pentadione
   (i) benzone
   (j) benzoic acid
   (k) sodiumethoxide
   (l) 3-methyl potassium butoxide
   (m) 2-3-diiodopentane
   (n) 2-fluoro 3-chloro 4-bromo 5-iodoheptane
   (o) paradichlorobenzene
   (p) orthodichlorobenzene

Unit 6: Sample Test 2

1. (a-k) Because of the possibility of more than one acceptable answer for questions a-k, no answers are provided here. (l) See below.



a polyester or a condensation polymer

2. (a) 4 (f) 8
   (b) 7 (g) 15
   (c) 21 (h) 12
   (d) 17 (i) 25
   (e) 9 (j) 16

Unit 2 Crossword Puzzle

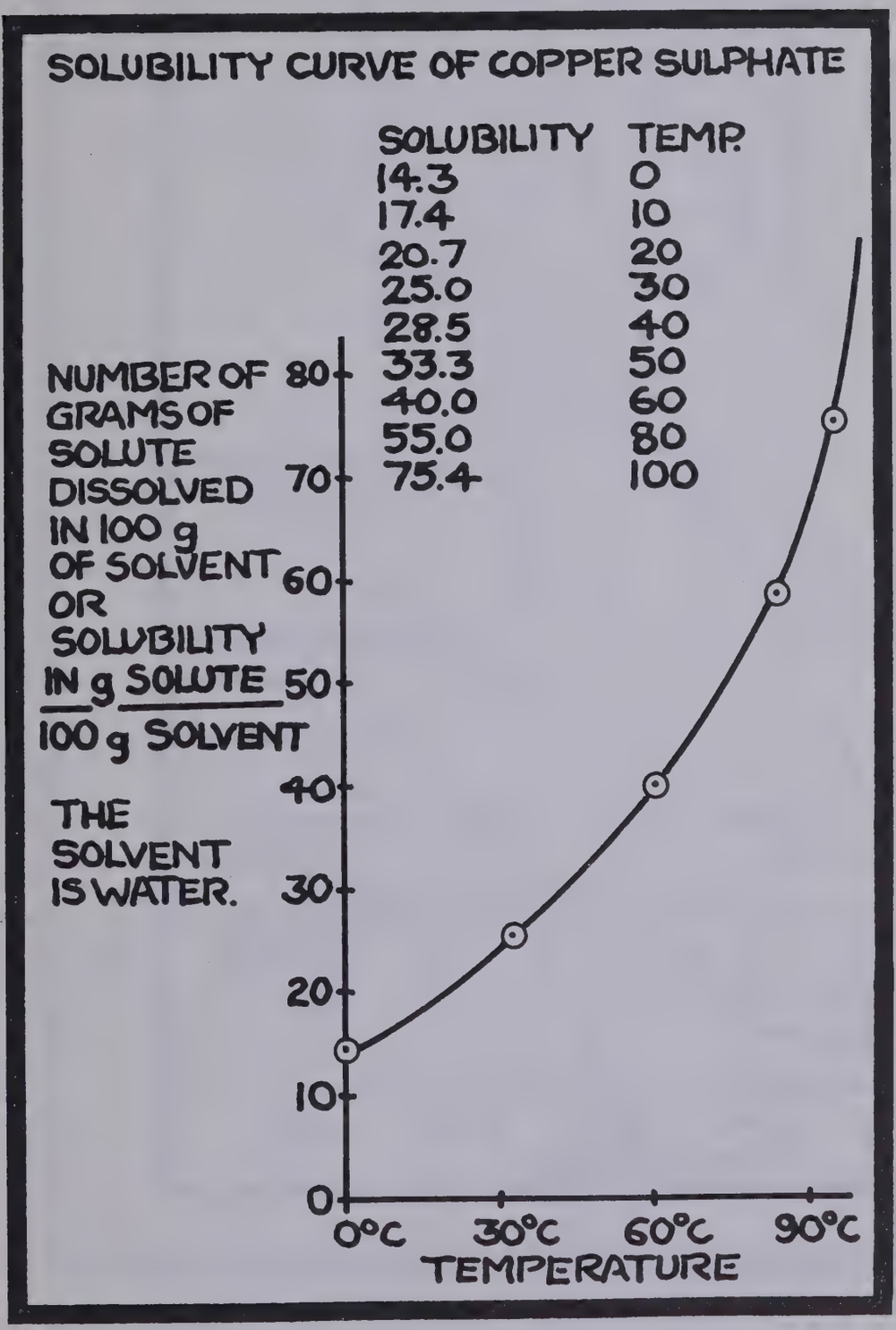
Down	Across
1. density	1. decomposition
2. combustion	5. regularity
3. mass	24. names
4. synthesis	25. In
5. reagent	26. atom
6. Ar	27. manganese
7. tin	28. test tube
8. molar	29. Dalton
9. Ni	30. pure
10. sulphur	31. mixture
11. equation	12. iron
12. indium	13. buoyant
13. Be	32. eudiometer
14. aluminum	33. Ir
15. pressure	34. temperature
16. temperature	35. atoms
17. kelvin	19. gas
18. symbol	36. homogeneous
19. Ga	37. Al
20. Os	38. Boyle
21. STP	39. ferrum
22. He	22. heat
23. Ne	

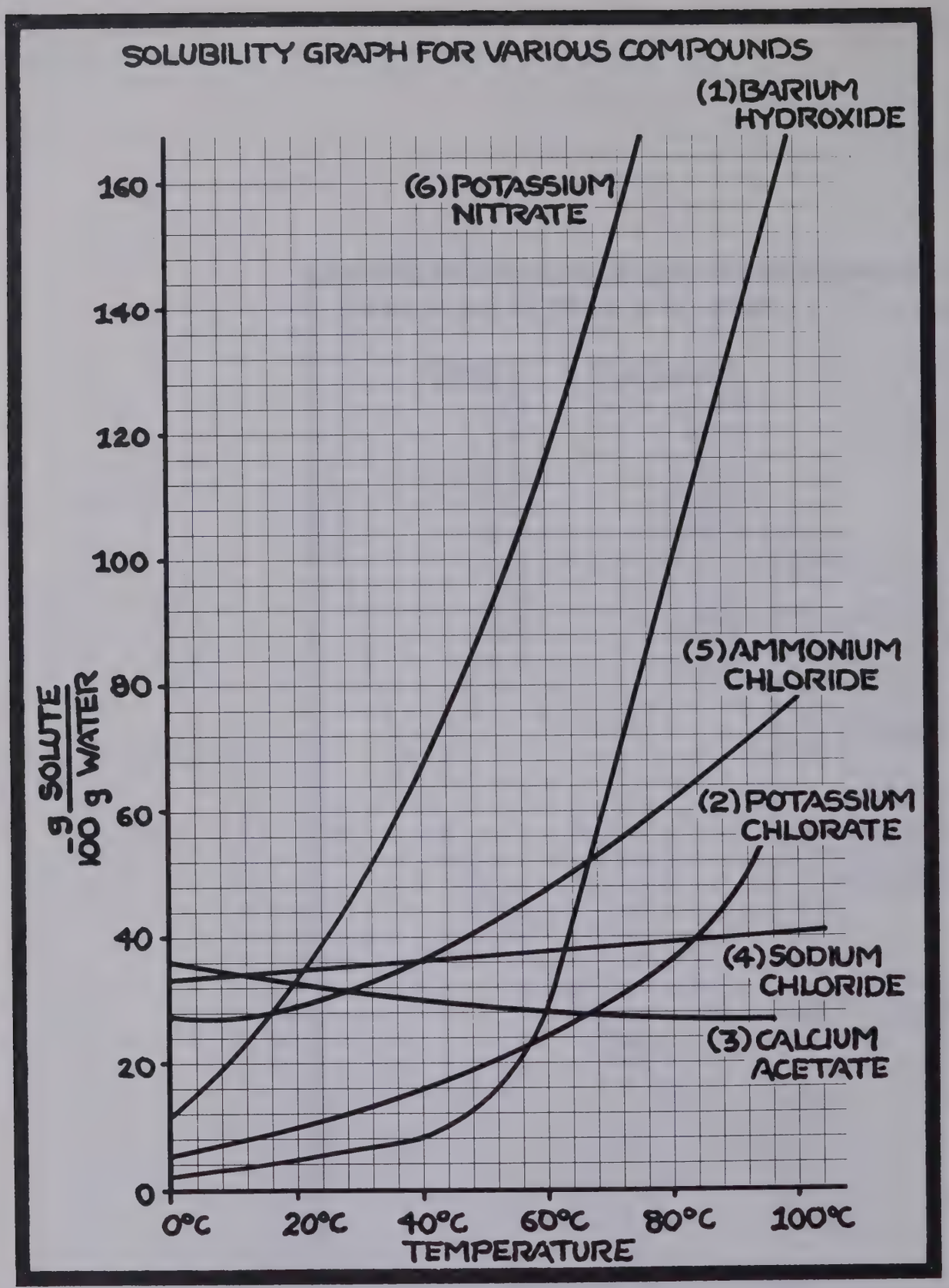
Unit 4 Crossword Puzzle

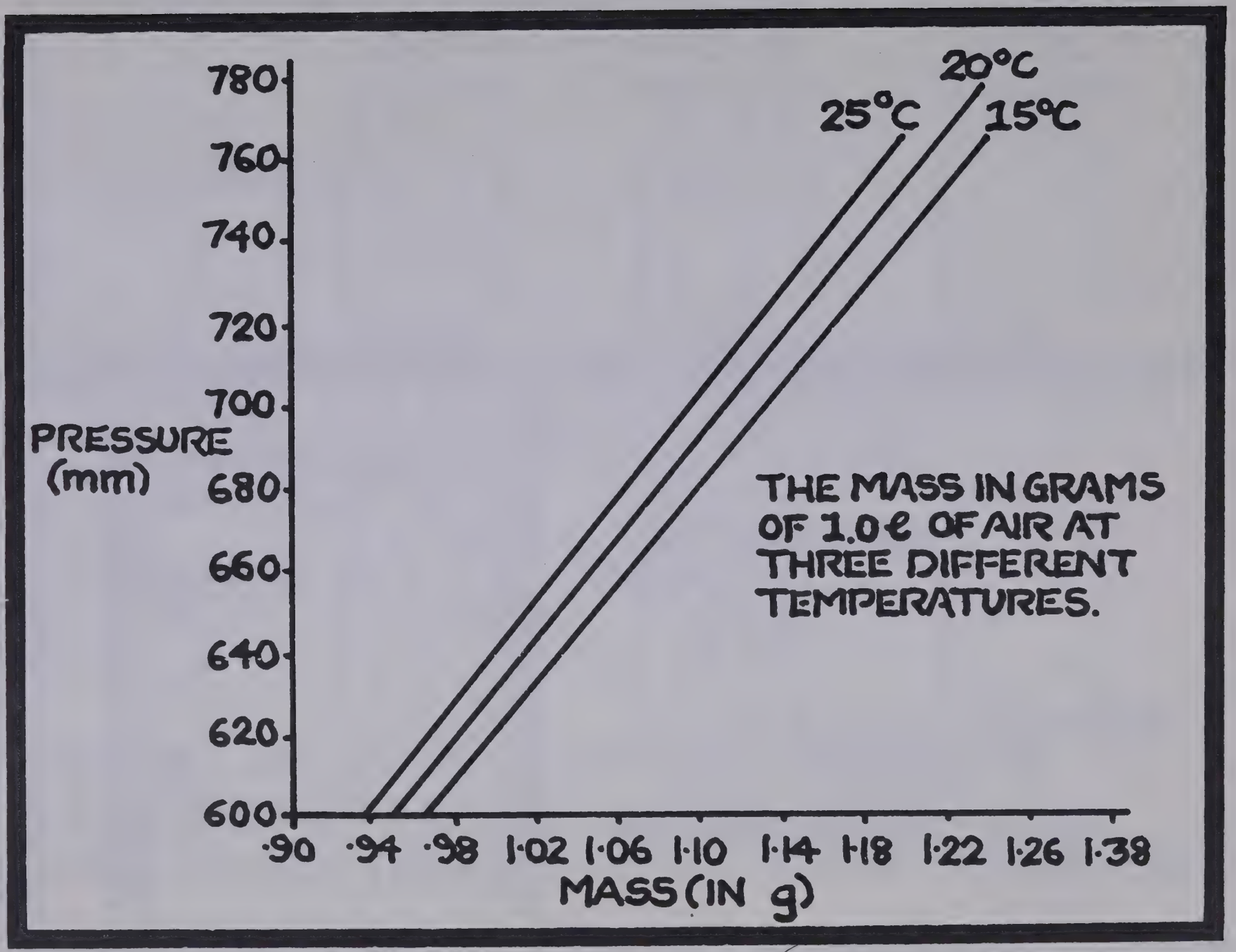
Across	Down
1. heat	32. oily
2. hypo	31. hydronium
3. anode	27. hot
4. tea	33. catalysis
5. gas	34. ion
6. acid	35. code
7. prefix	36. base
8. BTB	37. actinium
9. ASA	38. end
10. STP	39. meniscus
11. limewater	40. radon
12. organic	41. tin
13. soap	42. fun
14. hydroxide	43. binary
15. titration	44. per
16. displacement	45. fuel
17. displacement	46. and
18. oxidation	47. pen
19. Nb	48. radical
20. neutral	49. cat
21. ion	50. rat
22. strong	51. catalyst
23. hydroxide	8. burette
24. mole	52. dots
25. formula	10. salt
26. mossy	53. oxalate
27. hydrogen	4. tart
28. carbonate	61. meniscus
29. equation	55. energy
30. litmus	
60. us	

Appendix

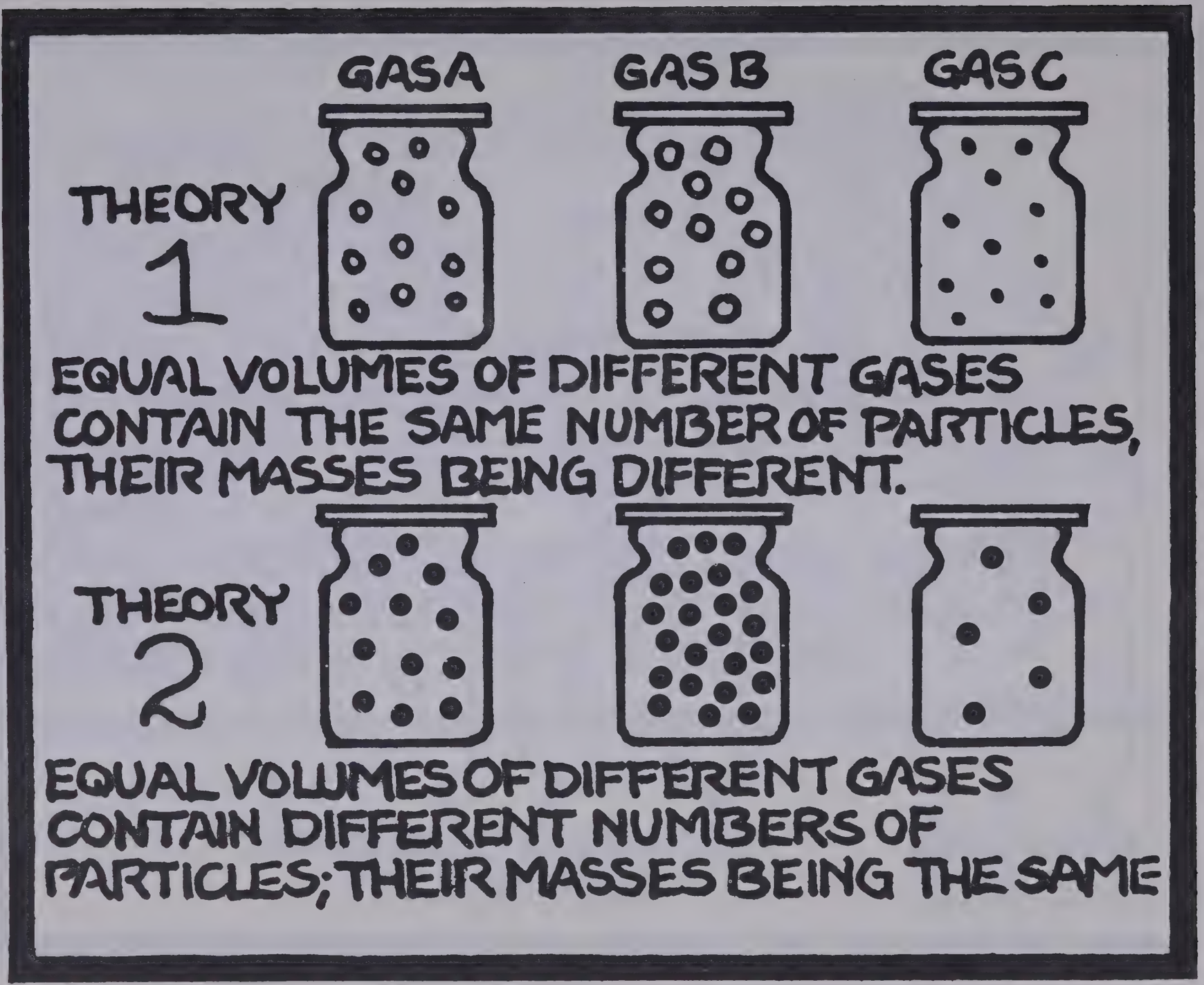
The following graphs and diagrams are reproduced from Units 1 through 6 of *The Nature of Matter* in a form suitable for conversion into overhead transparencies.

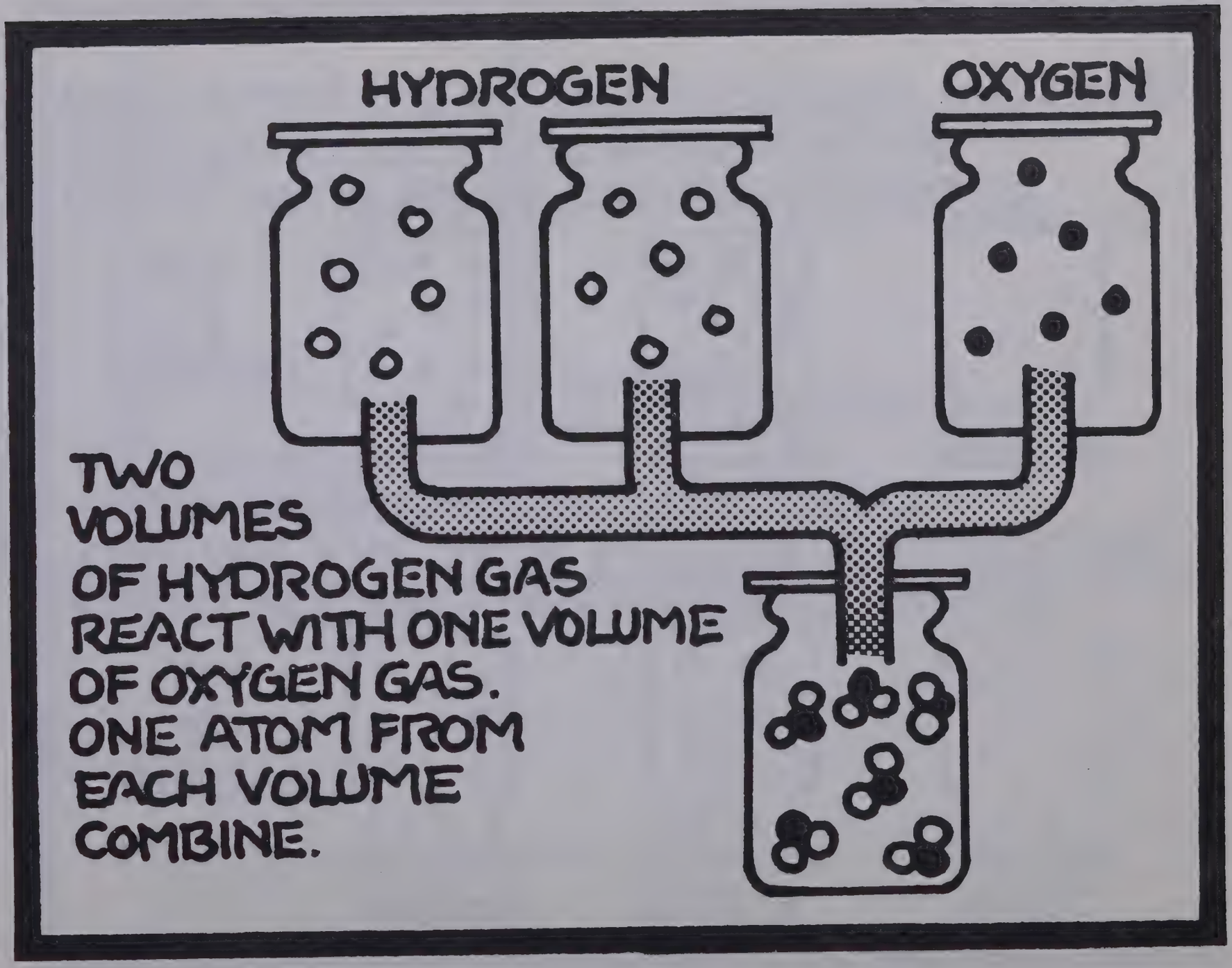


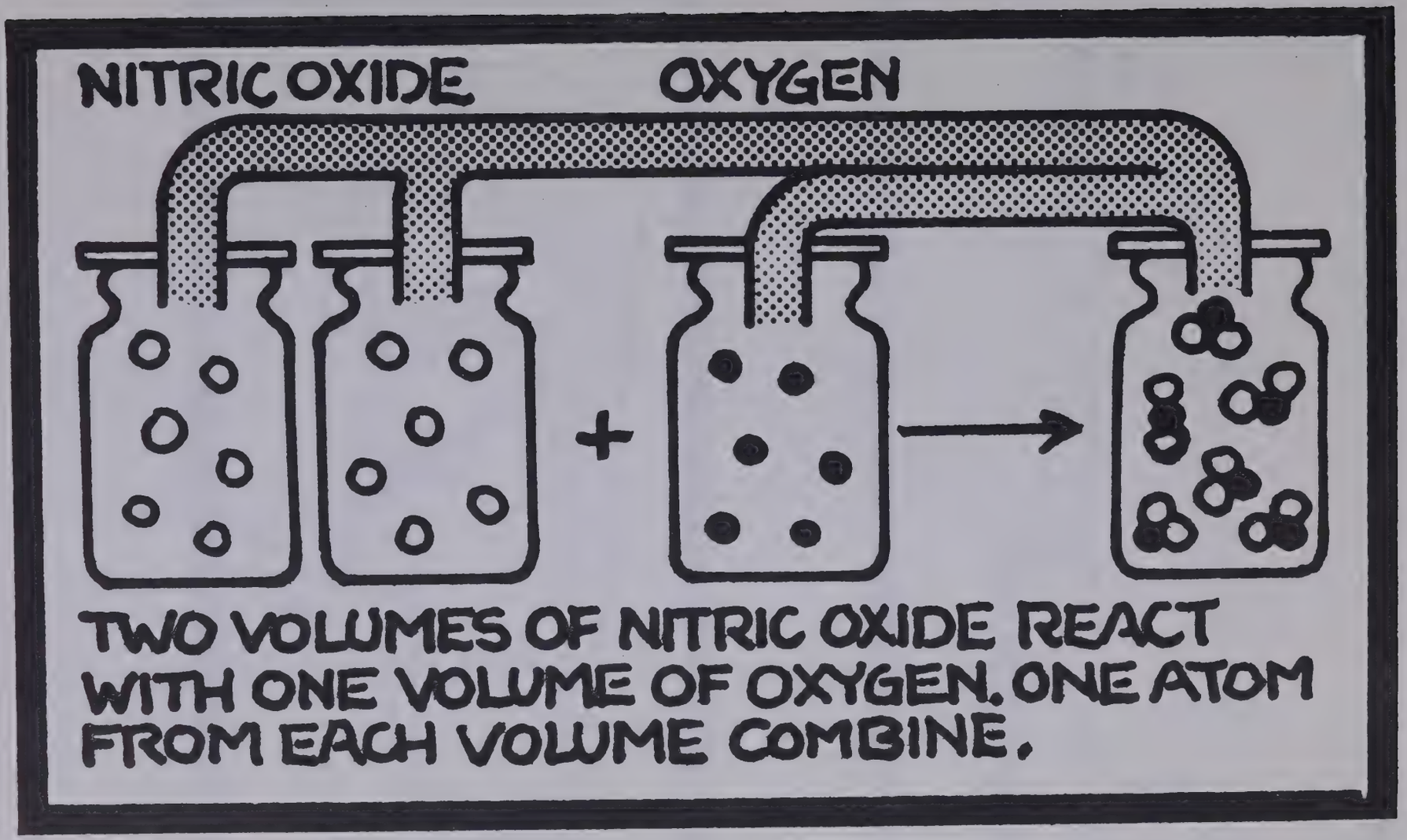

SOLUBILITY GRAPH FOR VARIOUS COMPOUNDS
(1) BARIUM HYDROXIDE
(6) POTASSIUM NITRATE
(5) AMMONIUM CHLORIDE
(2) POTASSIUM CHLORATE
(4) SODIUM CHLORIDE
(3) CALCIUM ACETATE
160
140
120
100
80
60
40
20
0
g SOLUTE / 100 g WATER
0°C
20°C
40°C
60°C
80°C
100°C
TEMPERATURE

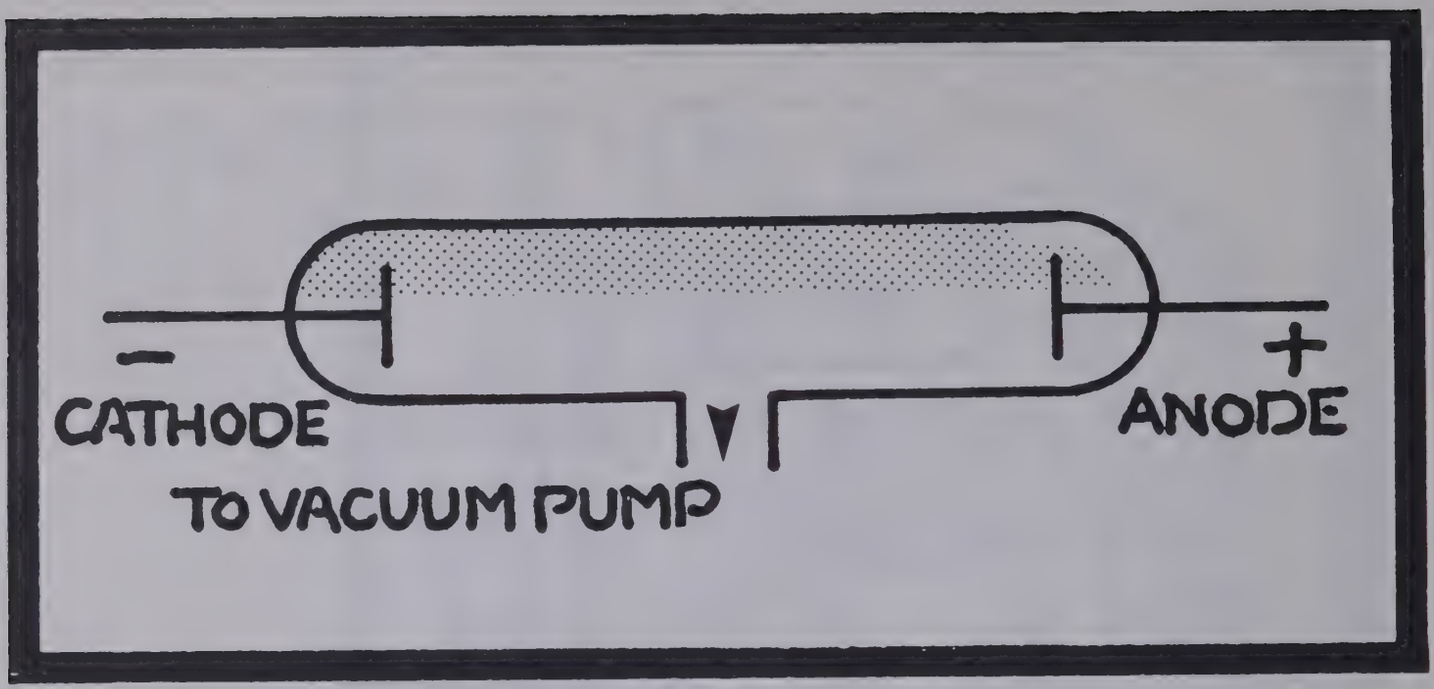

780
760
740
720
700
680
660
640
620
600
PRESSURE (mm)
20°C
25°C
15°C
THE MASS IN GRAMS OF 1.0 ℓ OF AIR AT THREE DIFFERENT TEMPERATURES.
·90 ·94 ·98 1·02 1·06 1·10 1·14 1·18 1·22 1·26 1·38
MASS (IN g)

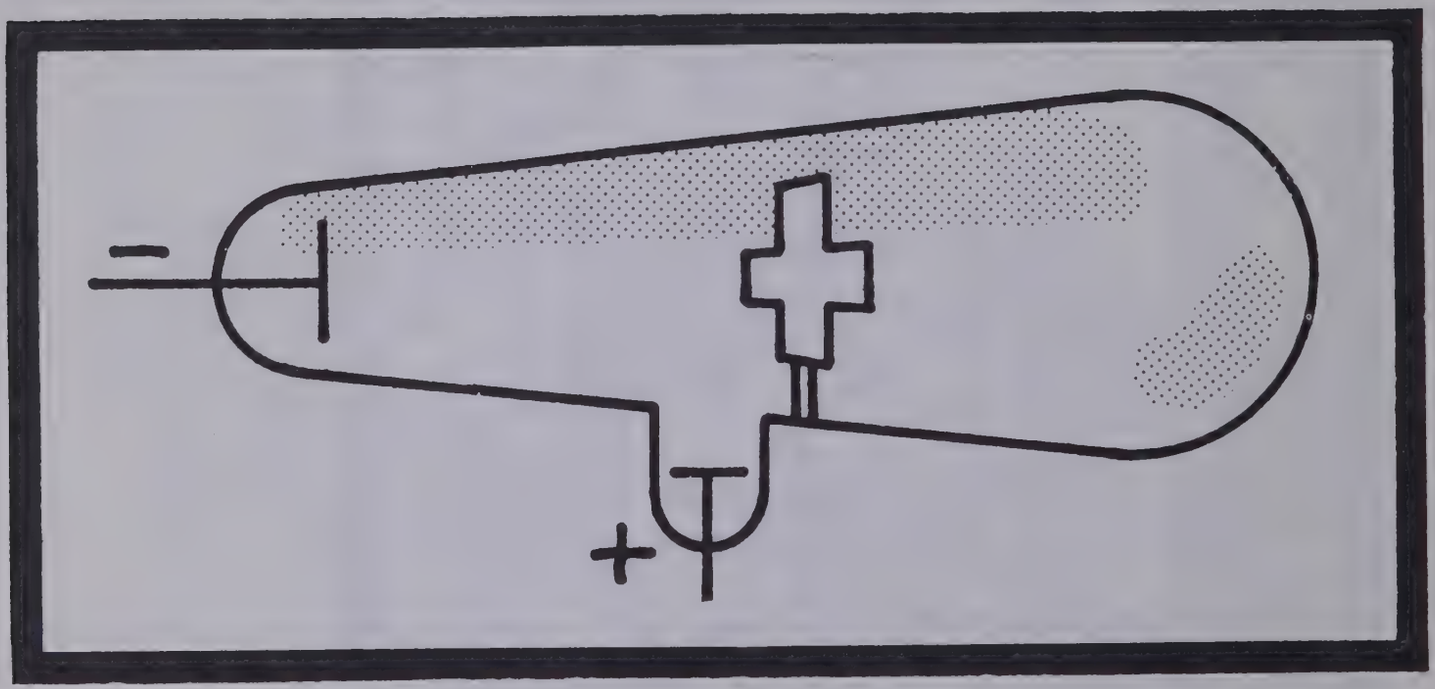
L
UPRIGHT
P=ATM.PRESSURE
+L
L
INVERTED
P=ATM.PRESSURE
-L
L
HORIZONTAL
P=ATM.PRESSURE

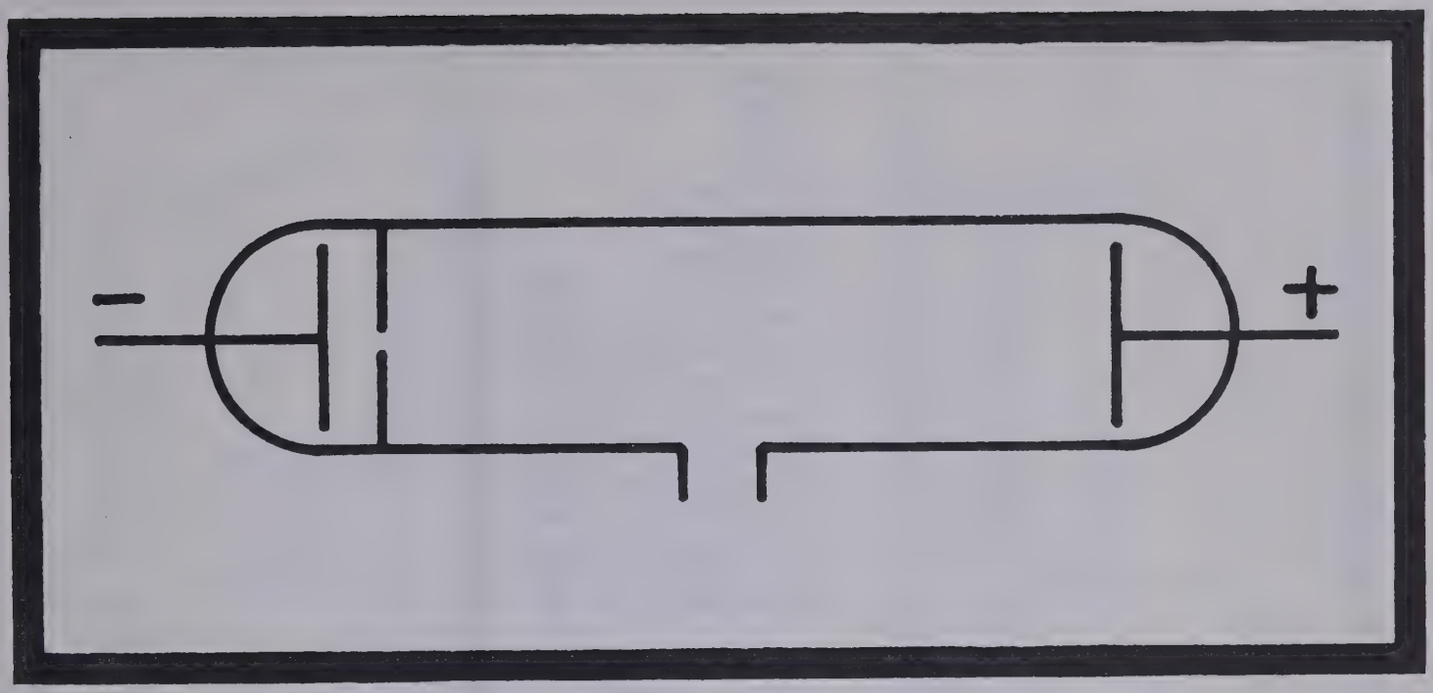

GAS A
GAS B
GAS C
THEORY 1
EQUAL VOLUMES OF DIFFERENT GASES CONTAIN THE SAME NUMBER OF PARTICLES, THEIR MASSES BEING DIFFERENT.
THEORY 2
EQUAL VOLUMES OF DIFFERENT GASES CONTAIN DIFFERENT NUMBERS OF PARTICLES; THEIR MASSES BEING THE SAME

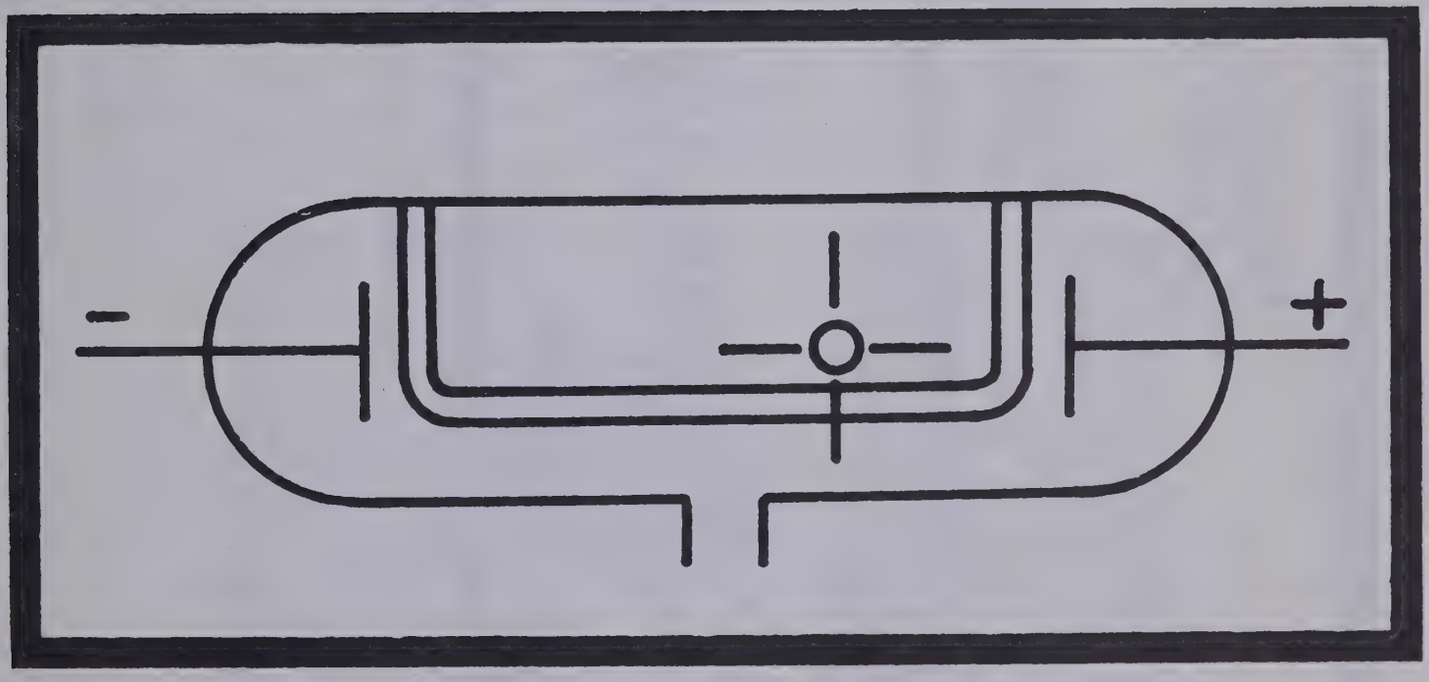

HYDROGEN
OXYGEN
TWO VOLUMES OF HYDROGEN GAS REACT WITH ONE VOLUME OF OXYGEN GAS. ONE ATOM FROM EACH VOLUME COMBINE.

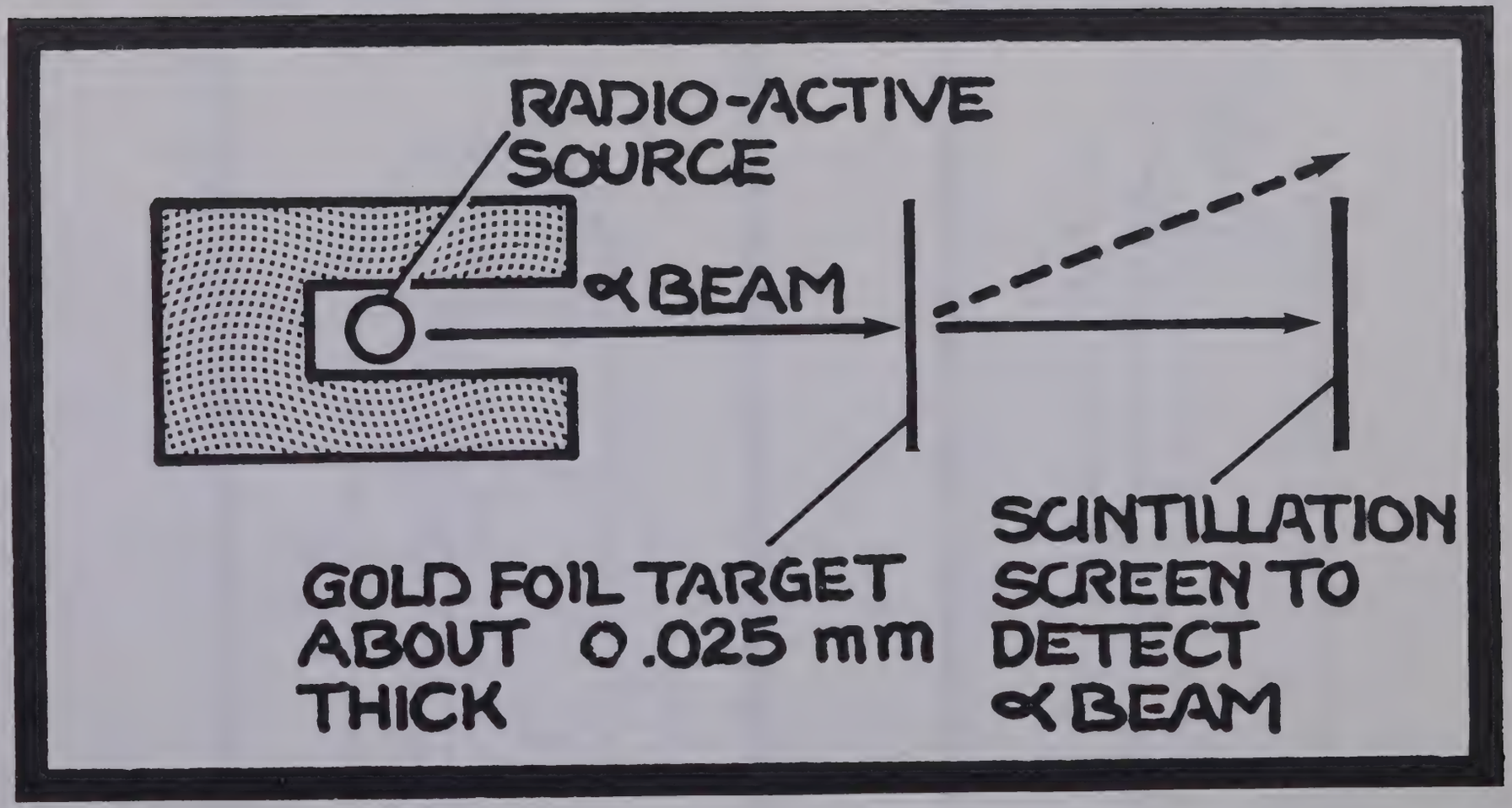

NITRIC OXIDE
OXYGEN
TWO VOLUMES OF NITRIC OXIDE REACT WITH ONE VOLUME OF OXYGEN. ONE ATOM FROM EACH VOLUME COMBINE.

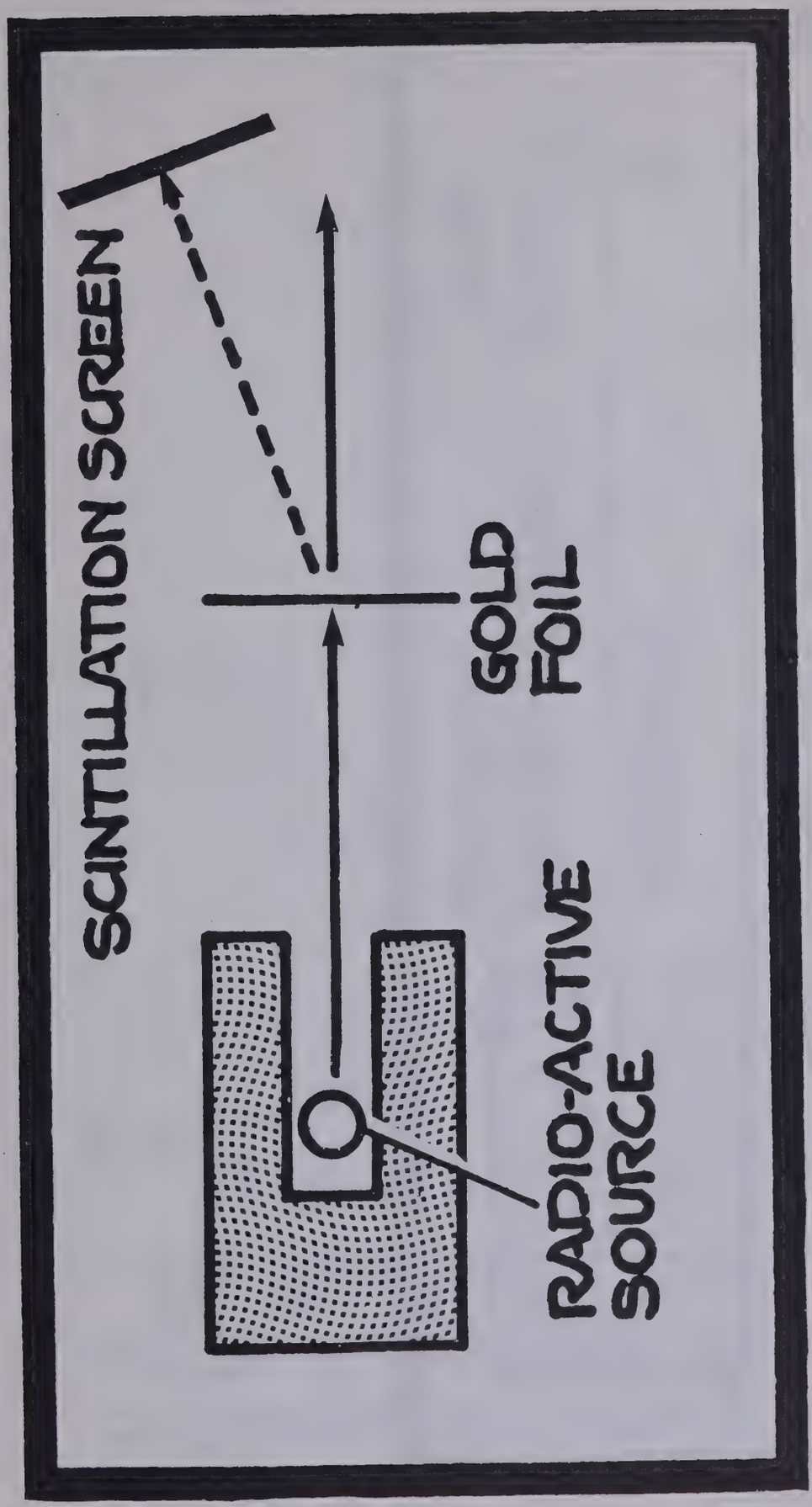

−
CATHODE
TO VACUUM PUMP
+
ANODE

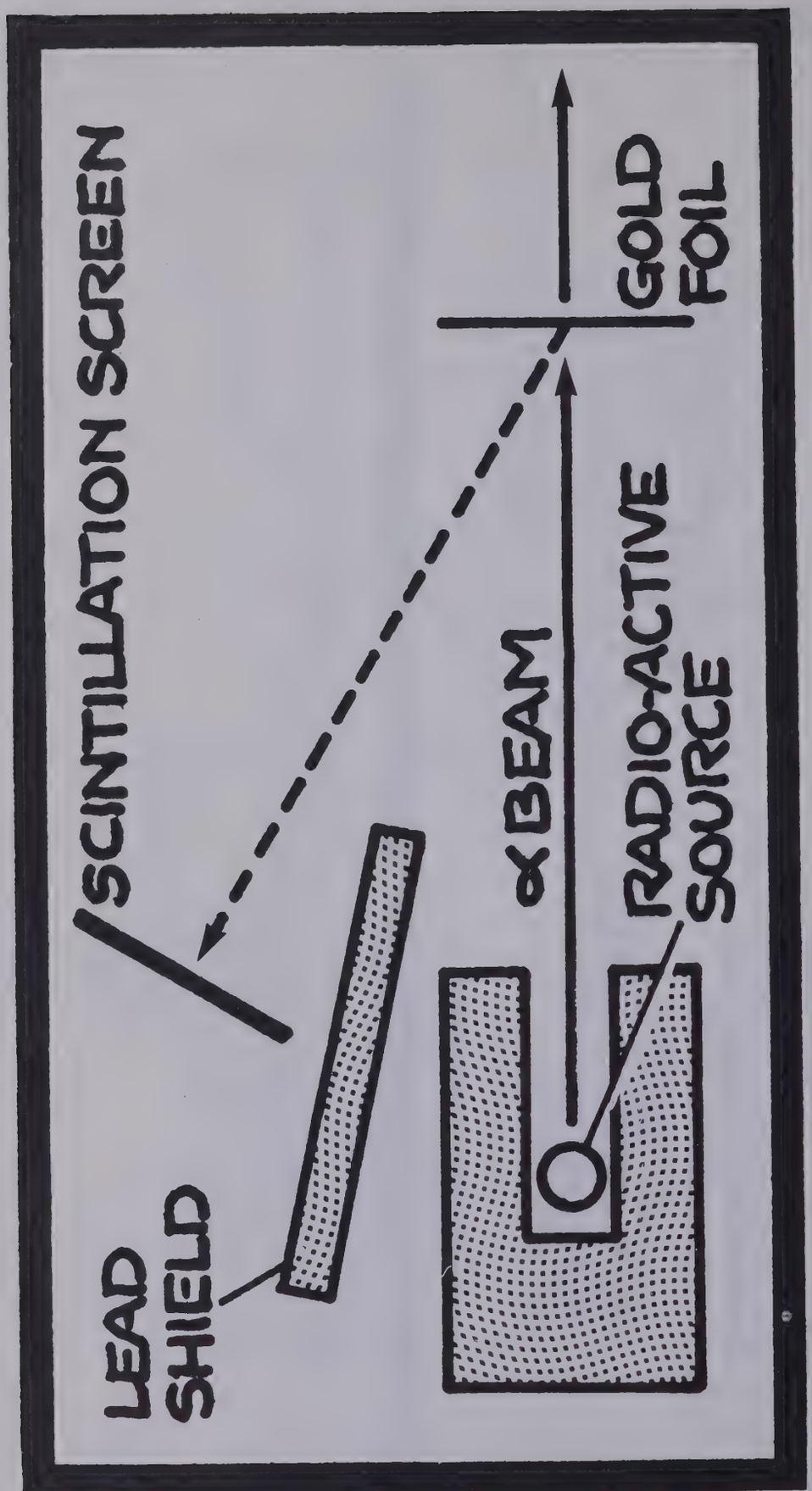

−
+



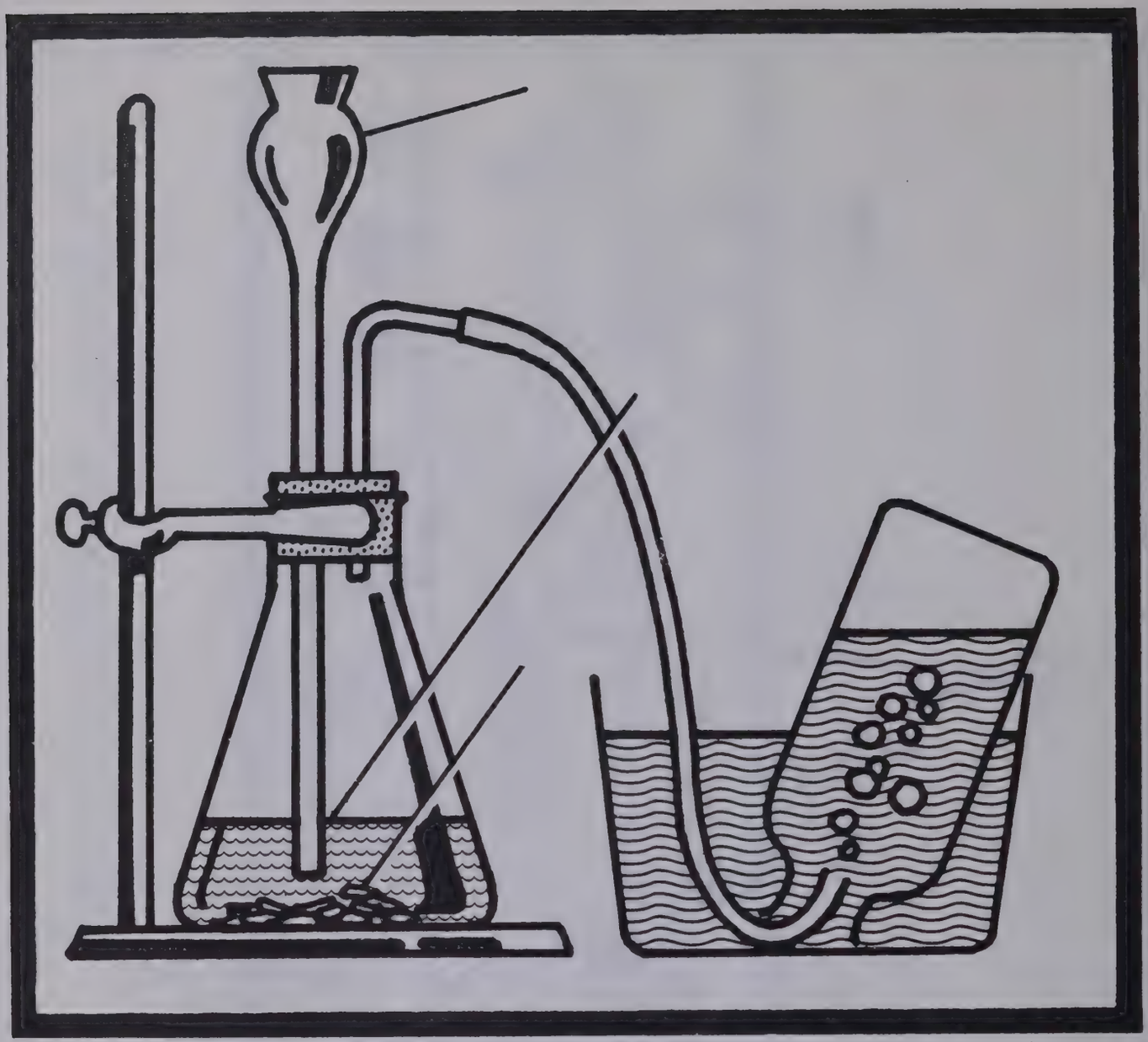


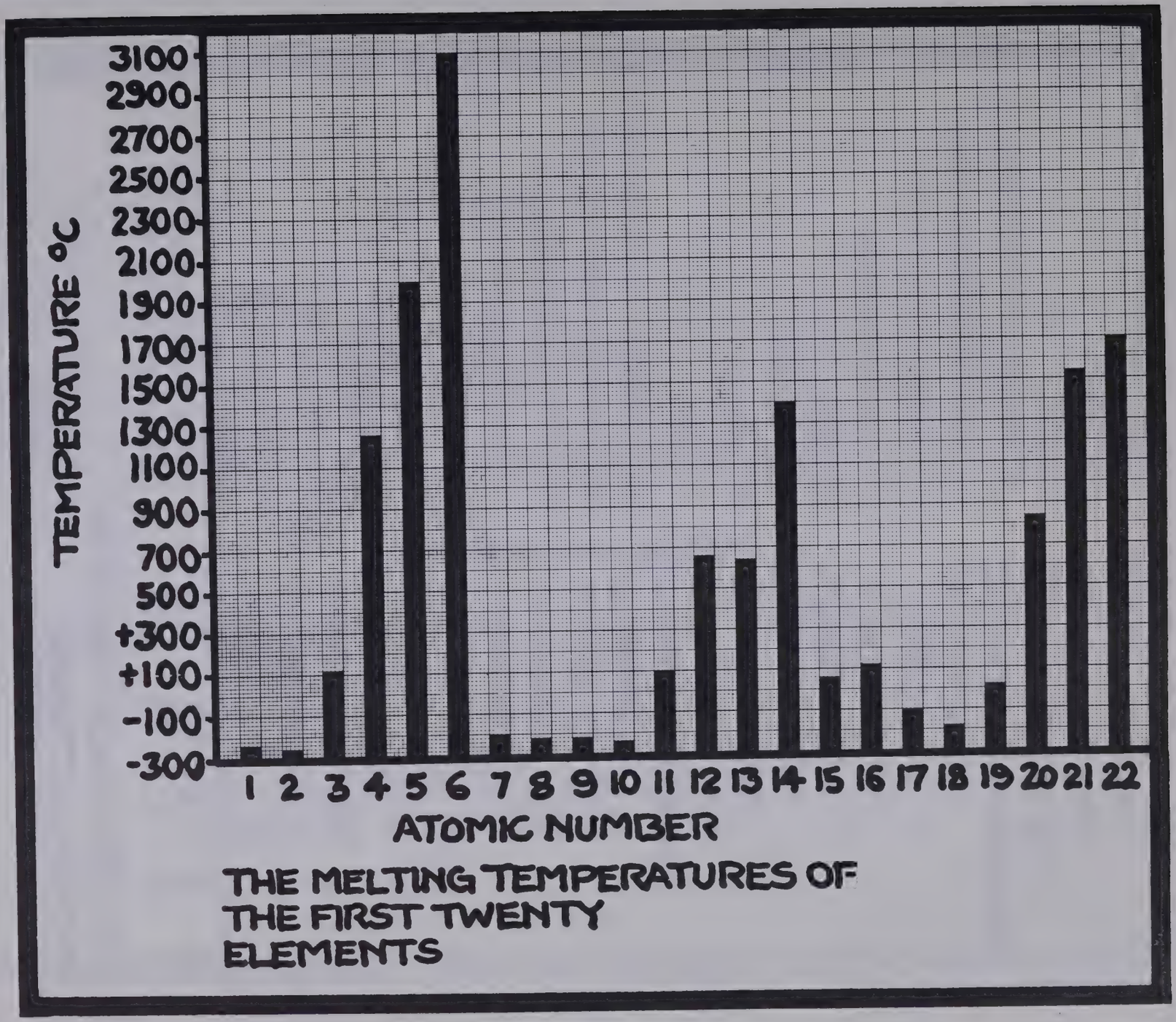

RADIO-ACTIVE SOURCE
α BEAM
GOLD FOIL TARGET ABOUT 0.025 mm THICK
SCINTILLATION SCREEN TO DETECT α BEAM


SCINTILLATION SCREEN
GOLD FOIL
RADIO-ACTIVE SOURCE


SCINTILLATION SCREEN
LEAD SHIELD
α BEAM
RADIO-ACTIVE SOURCE
GOLD FOIL

ENERGY (RELATIVE) TO REMOVE ONE ELECTRON
5
4
3
2
1
Z - ATOMIC NUMBER
2
4
6
8
10
12
14
16
18
20
H
He
Li
Be
B
C
N
O
F
Ne
Na
Mg
Al
Si
P
S
Cl
Ar
K
Ca

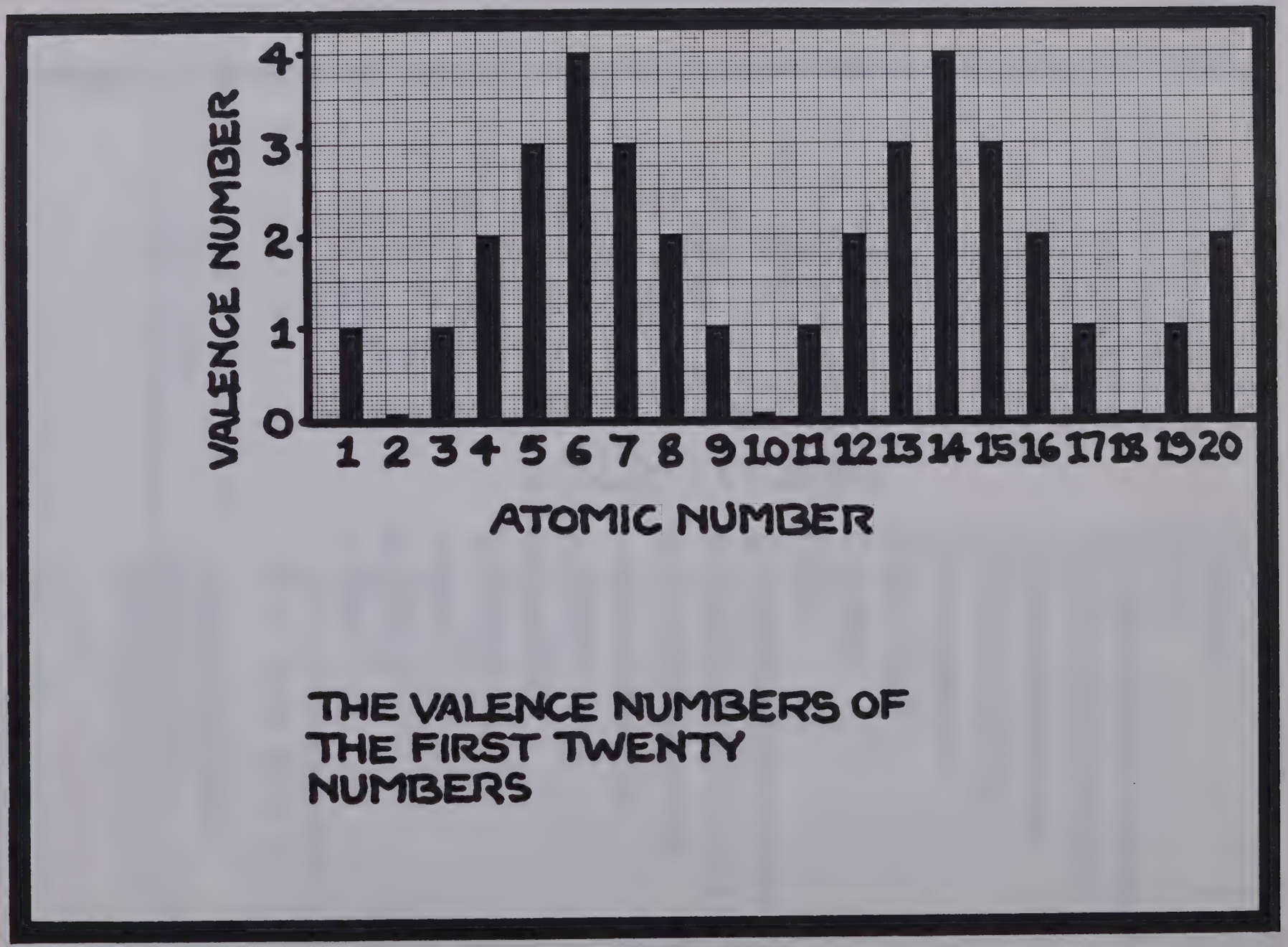




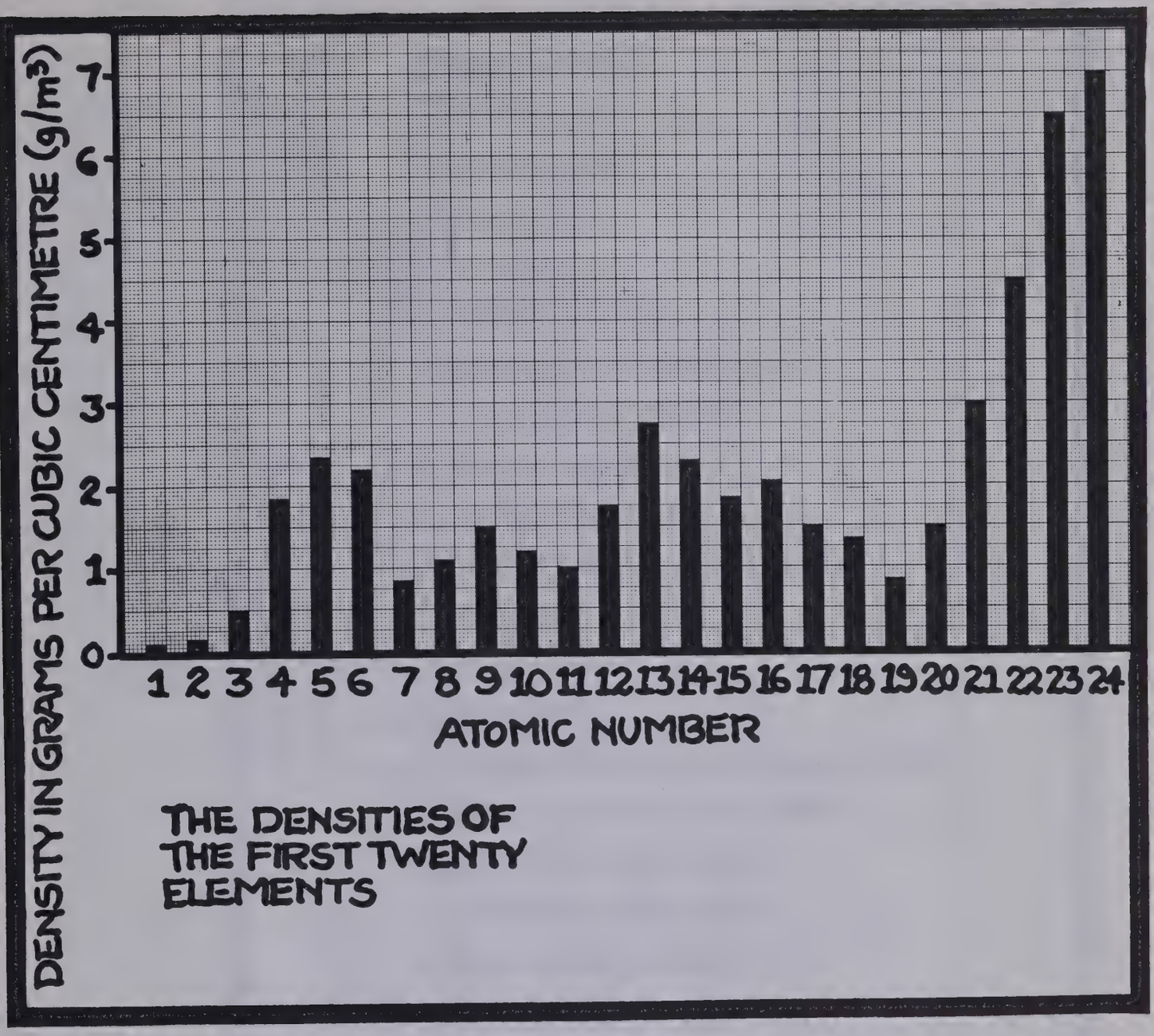
THE MELTING TEMPERATURES OF THE FIRST TWENTY ELEMENTS

THE COVALENT RADII OF THE FIRST TWENTY ELEMENTS

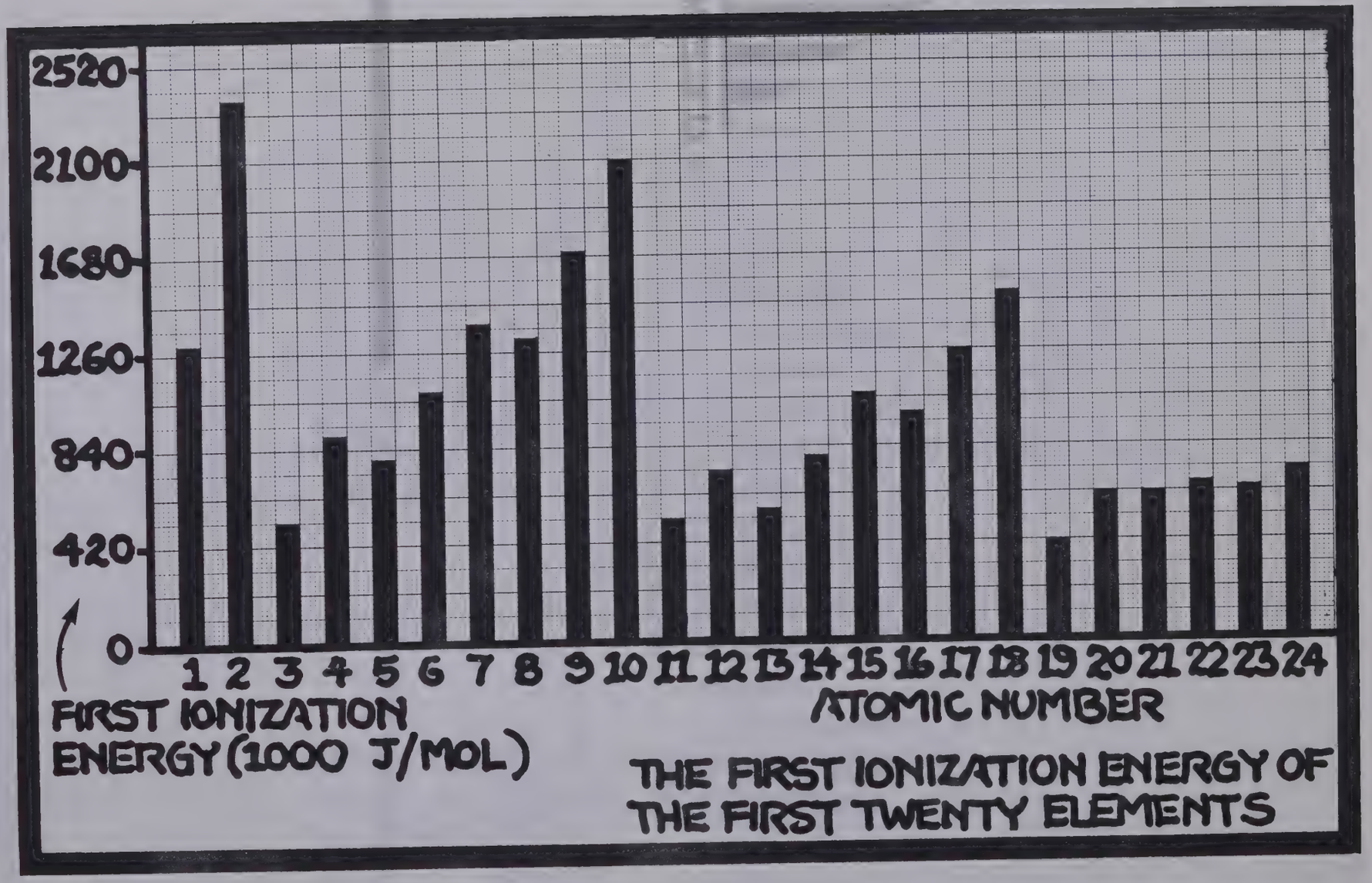
THE VALENCE SHELL ELECTRON POPULATION OF THE FIRST TWENTY ELEMENTS

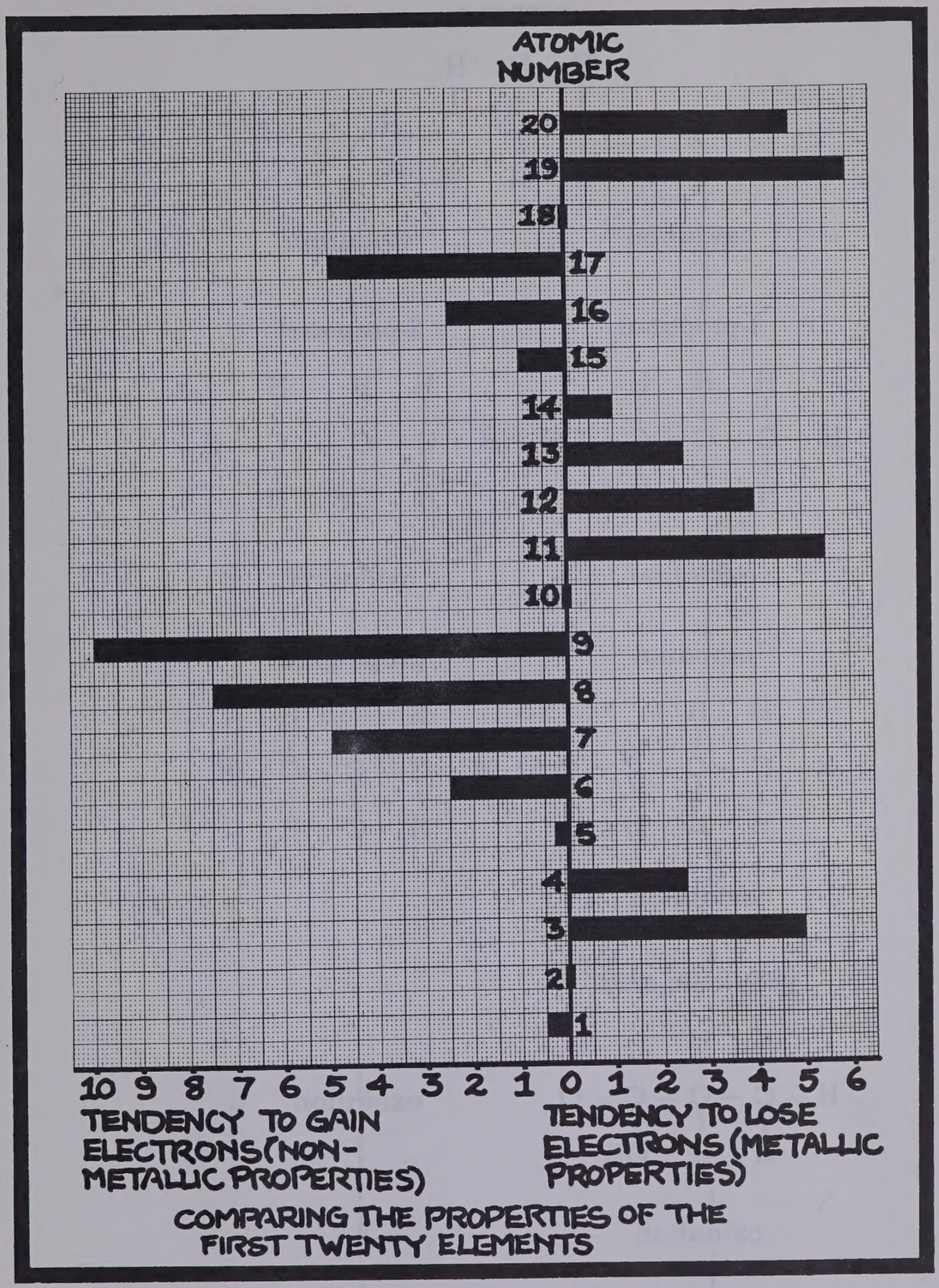


THE VALENCE NUMBERS OF THE FIRST TWENTY NUMBERS

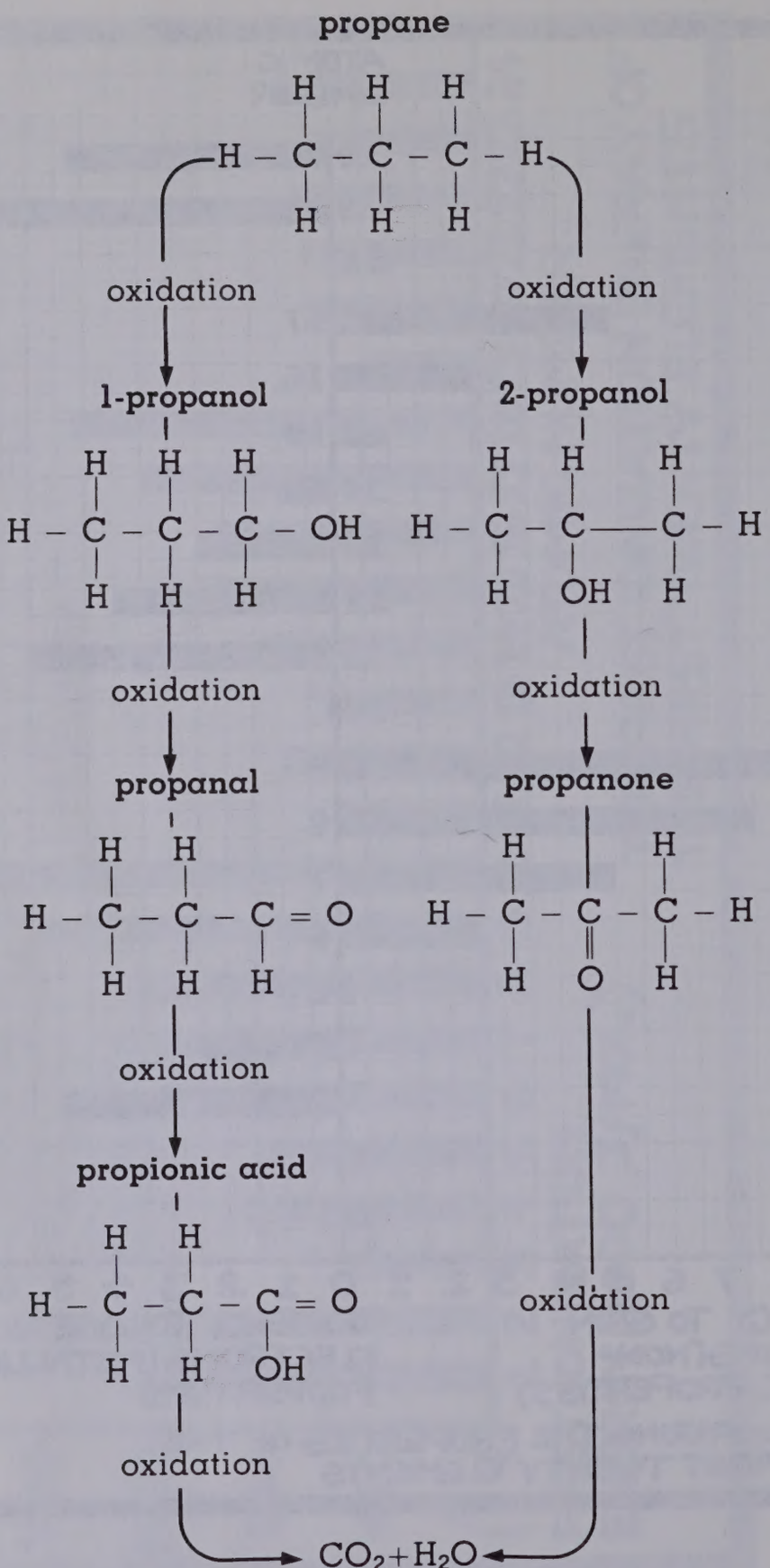


THE DENSITIES OF THE FIRST TWENTY ELEMENTS

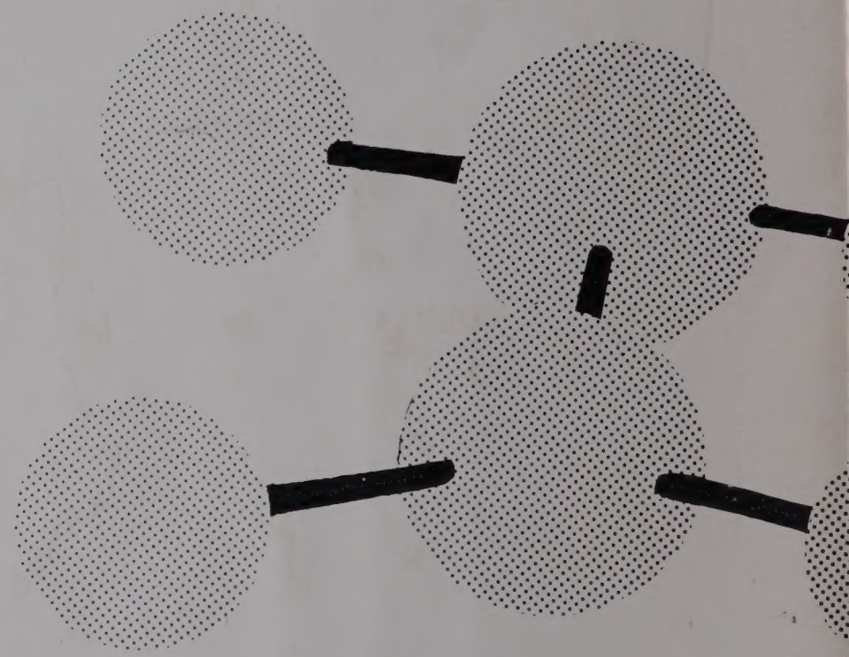

2520
2100
1680
1260
840
420
0
1 2 3 4 5 6 7 8 9 10 11 12 13 14 15 16 17 18 19 20 21 22 23 24
FIRST IONIZATION ENERGY (1000 J/MOL)
ATOMIC NUMBER
THE FIRST IONIZATION ENERGY OF THE FIRST TWENTY ELEMENTS



COMPARING THE PROPERTIES OF THE FIRST TWENTY ELEMENTS


propane
oxidation
oxidation
1-propanol
2-propanol
oxidation
oxidation
propanal
propanone
oxidation
propionic acid
oxidation
oxidation
CO_2+H_2O

B21769



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HEATH